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CXCIX.—HYDANTOINS

A NEW METHOD OF SYNTHESIZING N-ALKYL DERIVATIVES OF α-AMINO ACIDS METHYLTYROSINE

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[ELEVENTH PAPER]

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I. Surinamine

In 1824, Hüttenschmidt[1] and Overduin independently isolated from the bark of the cabbage tree (*Geoffroya surinamensis*) a nitrogenous substance which they called *surinamine* and *geoffroyine*, respectively. They assigned to the compound no empirical formula, but Hüttenschmidt states that it was soluble in water, formed crystalline salts with acids and was characterized by its property of giving a blue solution when dissolved in fuming nitric acid. This work of Hüttenschmidt's was repeated by Winck-

[1] Geiger's Magazin f. Pharm., 7, 287. Jahres. d. Chem., 1843, 322.

ler,[1] who apparently isolated the same base again from *Geoffroya surinamensis*, but he states that his product did not react with nitric acid, as mentioned by Hüttenschmidt, and attributed this behavior to an impurity, which was probably present in Hüttenschmidt's material. He confirmed, however, Hüttenschmidt's observation that the compound possesses basic properties and showed, furthermore, that it could be sublimed, but was largely decomposed by this treatment. Winckler also observed that the base was not precipitated from an aqueous solution by the addition of mercury chloride, silver nitrate, copper sulphate, zinc sulphate and tartar emetic.

Wittstein[2] later searched for this same base in an American ratanhia extract and succeeded in isolating a definite crystalline compound. He concluded from his analytical determinations and the chemical behavior of his substance that he was dealing with the amino acid *tyrosine*.

About 1862, Ruge,[3] under the direction of Städeler, took up the investigation of American ratanhia extract and isolated a crystalline substance, which was apparently identical with Wittstein's base. Ruge observed, as did Wittstein, that the compound reacted with mercury nitrate, in an aqueous solution, giving Hofmann's[4] reaction for tyrosine (Millon's test) and when heated with dilute nitric acid gave a beautiful red solution. He also showed that its salts readily underwent hydrolysis in water and dilute alcohol. Ruge concluded, however, that the base was not tyrosine and obtained analytical results which agreed well with the calculated values for a higher homologue of this amino acid—$C_{10}H_{13}O_3N$. He called this substance *ratanhine*. Wittstein[5] repeated Ruge's work and in several papers described the isolation of his base from ratanhia extract, but still maintained that he was not

[1] Jahrbuch d. Pharm., **2**, 159. Pharm. Centralb., **1840**, 120. Jahres. d. Chem., **1842**, 323.

[2] Vierteljahrsschr. pr. Pharm., **3**, 348, 485. Jahrbuch d. Pharm., **2**, 227. Jahres. d. Chem., **1854**, 656.

[3] Vierteljahrsschr. d. naturforsch. Ges. (Zürich), **6**, Heft 3. Jahres. d. Chem., **1862**, 493.

[4] Ann. Chem. (Liebig), **87**, 123.

[5] Vierteljahrsschr. pr. Pharm., **13**, 265. Chem. Centralb., **1864**, 1054. Jahres. d. Chem., **1864**, 613. Bull. soc. chim., **1865**, 444.

dealing with Ruge's *ratanhine* but with tyrosine. This work was followed by another paper by Ruge,[1] in 1865, in which he concludes, after a re-examination of his *ratanhine* and confirmation of his previous observations, that Wittstein's interpretation was wrong and that his base was undoubtedly *ratanhine* or a homologue of tyrosine.

In 1868, there appeared a short paper by Peckolt[2] in which he described the isolation of a basic substance from the bark of a species of leguminous trees from Brazil (*Ferreira spectabilis*). He named this substance *angeline*. For our knowledge of its chemical properties, however, we are indebted to Gintl,[3] who obtained it from the same source as Peckolt and subjected it to a very careful investigation. He concluded that the base was identical with Ruge's *ratanhine*. He described a large number of its salts, and also observed that it was precipitated by Nessler's reagent and partially by phosphotungstic acid. Gintl did not, however, establish its structure, but concluded from its characteristic behavior towards acids and bases that it was probably an amino acid and contained also a free hydroxyl group. He also examined its behavior on heating and observed that while it was stable below 150°, at a much higher temperature it melted with decomposition and was converted into a viscous liquid, which solidified on cooling. Gintl, however, did not examine this product, but states that it obviously was not the original ratanhine.

These observations of Ruge's and Gintl's were later confirmed by Kreitmair,[4] working under the direction of von Gorup-Besanez. He isolated a crystalline substance from ratanhia extract of a new source and observed that it behaved in every respect like Ruge's *ratanhine* and Peckolt's and Gintl's *angeline*. Kreitmair, however, did not succeed in isolating the base from all samples of ratanhia extract at

[1] J. prakt. Chem., **96**, 106. Chem. Centralb., **1865**, 1158.

[2] Z. österreich. Apothekervereins, **1868**, 518. Vierteljahrsschr. pr. Pharm., **18**, 444. Jahres. d. Chem., **1869**, 773.

[3] Ber. Wien. Akad., **58**, 443; **60**, 668. Z. österreich. Apothekervereins, **1868**, 518. Vierteljahrsschr. pr. Pharm., **18**, 444. Z. Chem., **1870**, 702. Bull. soc. chim., [2] **13**, 548. Jahres. d. Chem., **1869**, 774. J. prakt. Chem., **108**, 116. Chem. Centralb., **1869**, 394; **1870**, 281.

[4] Ann. Chem. (Liebig), **176**, 64.

his disposal and suggests that this was possibly due to the fact that the base, in these cases, was not in a free condition but linked in a protein like tyrosine.

Apparently no further attention was devoted to this alkaloid until 1893, when Hiller-Bombien[1] examined the bark of a *Geoffroya* according to Hüttenschmidt's method and isolated again the base *geoffroyine*.[2] He assigned to it the same formula, $C_{10}H_{13}O_3N$, as was originally suggested by Ruge and decided that it was identical with *ratanhine* and *angeline*. He also isolated the same base from the bark of *Andira anthelmintica* and assigned to it a new name, *andirine*. Hiller states in this paper that the base is probably *methyltyrosine*.

The first evidence which suggested the molecular structure of *surinamine* was obtained by Blau[3] in 1908. He obtained his base from the bark of a *Geoffroya*, carefully purified it in the form of its hydrochloride, and found that the free base melted at 246° with decomposition. He subjected it to a destructive distillation and obtained a viscous, basic oil, which solidified on cooling. Enough of this, however, was not obtained for experimental purposes and analyses, but he succeeded in obtaining a characteristic platinum chloride salt. This gave, on analysis, results for platinum agreeing with the calculated value for the double platinum chloride salt of methyl-*p*-hydroxyphenylethylamine,

$$(HOC_6H_4CH_2CH_2NHCH_3.HCl)_2.PtCl_4.$$

The salt crystallized in needles and melted at 205°. That the alkaloid contained the nucleus $HO.C_6H_4$— was proved by the fact that *p*-hydroxybenzoic acid, $HO.C_6H_4COOH$, was formed when the base was heated with strong potash solution. Blau did not, however, succeed in showing that methylamine was formed by this treatment. The evidence, so far obtained, therefore indicates that *surinamine* has the formula $C_{10}H_{13}O_3N$ and is to be assigned the constitution of *monomethyltyrosine*,

$$p\text{-}HOC_6H_4CH_2CH(NHCH_3)COOH.$$

[1] Arch. f. Pharm., **230**, 513. Chem. Centralb., **1893**, I, 36.
[2] Overduin's nomenclature.
[3] Zeit. physiol. Chem., **58**, 153.

An amino acid having this structure has recently been synthesized by Friedmann and Gutmann,[1] but being unable to obtain the natural base for comparison they were not able to establish the identity of the two.

II. The Synthesis of Methyltyrosine (Surinamine?)

Two general methods are available for the synthesis of a monoalkyl derivative of an α-amino acid: (1) Heating an α-halogenated fatty acid (II) with a primary amine:

$$\underset{\text{II}}{\text{RCH(hal)COOH}} + \text{CH}_3\text{NH}_2 = \text{RCH(NHCH}_3\text{)COOH} + \text{H.hal}$$

and (2) condensation of the nitrile of an α-hydroxy acid (III) (cyanohydrin) with the amine, forming the nitrile of the amino acid (IV), and then the conversion of the nitrile into the acid by hydrolysis:

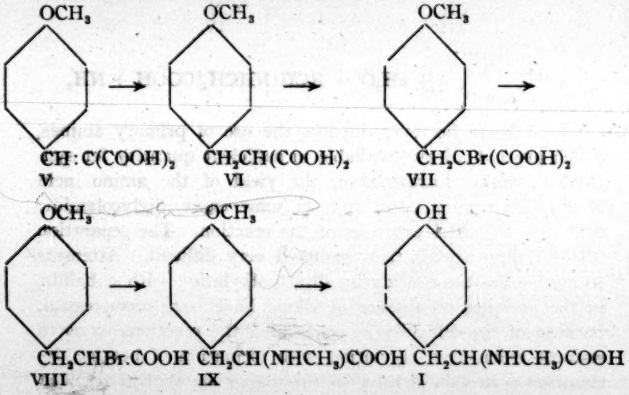
$$\underset{\text{III}}{\text{RCH}\begin{matrix}\text{OH}\\ \diagdown\\ \text{CN}\end{matrix}} + \text{CH}_3\text{NH}_2 = \underset{\text{IV}}{\text{RCH}\begin{matrix}\text{NHCH}_3\\ \diagdown\\ \text{CN}\end{matrix}} + \text{H}_2\text{O}$$

$$\text{RCH}\begin{matrix}\text{NHCH}_3\\ \diagdown\\ \text{CN}\end{matrix} + 2\text{H}_2\text{O} = \text{RCH(NHCH}_3\text{)COOH} + \text{NH}_3$$

Both methods, however, involve the use of primary amines, which are not always available in sufficient quantity for synthetical work. Furthermore, the yield of the amino acid is oftentimes very poor, and, in some cases, diglycolamidic acids are the chief products of the reaction. The separation of the amino acid is then rendered very difficult. Attempts to synthesize these acids by direct alkylation with a halide, in the presence or absence of alkali, have been unsuccessful, because of the difficulty in controlling the reactions so as to prevent exhaustive alkylation with formation of betaine derivatives. In this chapter of this paper we shall describe a new method of synthesizing monoalkylated amino acids which does not involve the use of a primary amine. It has been applied successfully in the synthesis of the interesting amino acid, *methyltyrosine*.

[1] Biochem. Z., **1910**, 490.

Friedmann and Gutmann's[1] method of synthesizing methyltyrosine, and also methylphenylalanine, is an application of Method (1) described above and therefore involves the use of methylamine. They prepared the acid in the following manner: Anisalmalonic acid (V) was first prepared by condensation of anisaldehyde with malonic acid,[2] and then converted into *p*-methoxybenzylmalonic acid (VI) by reduction with sodium amalgam. By bromination in ether solution the corresponding bromo derivative (VII) was formed, and this then transformed into α-bromo-β-*p*-methoxyphenylpropionic acid (VIII) by heating at 130°. This was then converted into α-methylamino-β-*p*-methoxyphenylpropionic acid (IX) by heating with 33 per cent. methylamine solution and the amino acid then transformed into methyltyrosine (I) by heating with hydrobromic acid. In other words, their method of synthesis involves six different operations, which may be represented by the following formulas:



Their yield of methyltyrosine, calculated from the weight of anisalmalonic acid used, was about 28 per cent. of the theoretical.

[1] *Loc. cit.*

[2] Knoevenagel: Ber. d. chem. Ges., **31**, 2607.

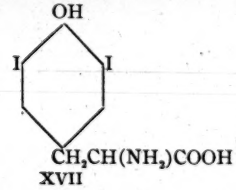
We now find that methyltyrosine can be synthesized easily from anisalhydantoin (X) by a method which involves only three operations. The anisalhydantoin was prepared by condensing anisaldehyde with hydantoin, and has already been described in a previous publication from this laboratory.[1] We find that this hydantoin (X) reacts with methyl iodide, in dilute alcohol and in the presence of alkali, giving a yield of 1,3-dimethylanisalhydantoin (XII) corresponding to about 87 per cent. of the theoretical. 1-Methylanisalhydantoin (XI) is the primary product of this reaction. When the dimethylhydantoin (XII) is digested with hydriodic acid, methyl iodide is evolved and it undergoes reduction at the double bond, forming 1,3-dimethyltyrosinehydantoin (XIV). This is then hydrolyzed with barium hydroxide, when methylamine is evolved and the barium salt of methyltyrosine (I) is obtained. After removing the barium in the usual manner, we obtained a yield of methyltyrosine which corresponded to about 32 per cent. of the theoretical, when calculated from the weight of anisalhydantoin (X) used. The method of synthesis can also be modified as follows: 1,3-Dimethylanisalhydantoin (XII) is first reduced by tin and hydrochloric acid to 1,3-dimethylanisylhydantoin (XV) and the latter then converted, by hydrolysis with barium hydroxide, into α-methylamino-β[L]*p*–methoxyphenylpropionic acid (XVI). This acid was described by Friedmann and Gutmann,[2] who showed that it can be transformed into methyltyrosine (I) by heating with hydrobromic acid. We accomplished the same result by boiling the acid with hydriodic acid and obtained a yield of methyltyrosine corresponding to 65 per cent. of the theoretical. 1-Methylanisalhydantoin (XI) is reduced by hydriodic acid, forming 1-methyltyrosinehydantoin (XIII) with evolution of methyl iodide. These various changes are represented by the following formulas:

[1] Wheeler and Hoffman: THIS JOURNAL, **48**, 368.
[2] *Loc. cit.*

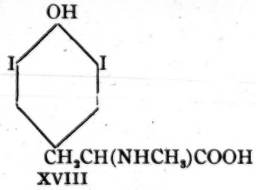
Friedmann and Gutmann state that their methyltyrosine decomposed sharply at 318°. It has been our experience that this decomposition point is very indefinite. Our acid could be decomposed at as low a temperature as 265° if heated for a long time, and when the bath was heated rapidly the decomposition point could be raised to 320°. Gintl[1] states that *surinamine* melts at 246°, while Pictet and Biddle[2] give 257°. Our acid agreed in all its properties, so far as we were able to judge, with the *methyltyrosine* synthesized by Friedmann and Gutmann.

III. Action of Iodine on Methyltyrosine

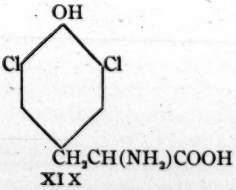
The naturally occurring halogen amino acid—*iodogorgoic acid*—is 3,5-diiodotyrosine (XVII) and was first synthesized in this laboratory.[3] It was prepared by the action of iodine on tyrosine and its constitution was later established by Wheeler and Johns.[4] The corresponding dibromo and dichloro acids (XIX) and (XX), have also been synthesized.[5] No



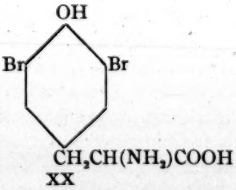
XVII



XVIII



XIX



XX

[1] *Loc. cit.*

[2] Vegetable Alkaloids, p. 463.

[3] Wheeler and Jamieson: THIS JOURNAL, **33**, 365; **38**, 356.

[4] THIS JOURNAL, **43**, 11.

[5] Von Gorup: Ann. Chem. (Liebig), **128**, 281. Johnson and Hoffman: THIS JOURNAL, **47**, 20. Wheeler, Hoffman and Johnson: J. Biol. Chem., **10**, 147.

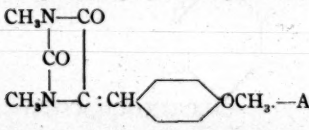
organic compound of known structure containing chlorine or bromine, however, has been isolated from animal tissues.

We have now prepared the higher homologue of iodogorgoic acid. We find that iodine reacts with our methyltyrosine in the same manner as with tyrosine, forming, apparently, the corresponding 3,5-diiodoamino acid, viz., α-methylamino-β-(3,5-diiodo-4-hydroxyphenyl)propionic acid or *methyliodogorgoic acid* (XVIII). Our assumption that iodine substitutes in positions analogous to those taken in tyrosine is confirmed by the fact that the halogen acid gives no red color when warmed with Millon's reagent. The acid is less soluble in water than iodogorgoic acid and resembles, in its chemical behavior, very closely this natural halogen acid. A description of the acid is given in the experimental part of this paper.

EXPERIMENTAL PART

The anisalhydantoin which was used in this investigation was prepared by the condensation of hydantoin with *p*-anisaldehyde.[1]

1,3-Dimethyl-4-anisalhydantoin,



—Anisalhydantoin and 2.2 molecular proportions of potassium hydroxide were dissolved in about 12 parts of 50 per cent. alcohol. An excess of methyl iodide (2.3 mols.) was then added, and the solution carefully heated on a steam bath for about 3 hours or until it was neutral to litmus. The solution was then concentrated to expel the alcohol, the residue triturated with cold, dilute alkali to remove inorganic salts and any incompletely methylated hydantoin, and the dimethylhydantoin then extracted with ether. The ether solution was then washed with a little water, dried, and the solvent evaporated, when the hydantoin was obtained in the form of yellow, prismatic needles. The compound was purified by recrystallization from ether or alcohol

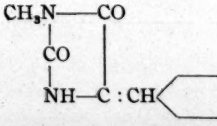
[1] Wheeler and Hoffman: *Loc. cit.*

and separated in prisms, which melted at 84°–85° to a yellow oil. This hydantoin is very soluble in hot alcohol, readily soluble in cold alcohol and ether, and practically insoluble in cold water. The yield of hydantoin was 70 per cent. of the theoretical. Analysis (Kjeldahl):

	Calculated for $C_{12}H_{14}O_3N_2$	Found
N	11.37	11.04

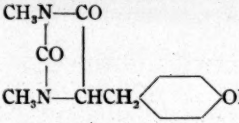
On acidifying the water solution above, after the extraction of the dimethylhydantoin with ether, a quantity of 1-methyl-4-anisalhydantoin (see below) separated, corresponding to about 15–20 per cent. of the calculated amount. By methylating this again with an excess of methyl iodide the total yield of the dimethylhydantoin was finally raised to over 80 per cent. of the theoretical. In order to avoid this second methylation, attempts were made to effect a complete methylation by using a large excess of potash and methyl iodide. After several trials, however, it was found that the best yield was obtained when to the neutral solution, after methylation as described above, two-thirds of a molecular proportion of alkali and a slight excess of methyl iodide were added, and the solution warmed one hour longer. In this manner the yield of the dimethylhydantoin was raised to 87 per cent. of the calculated and only a trace of the 1-monomethylhydantoin was obtained.

1-Methyl-4-anisalhydantoin,

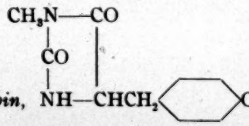


—This hydantoin, which was obtained as a by-product in the preceding experiment, was easily prepared by alkylation of anisalhydantoin with the required proportions of alkali and methyl iodide. It crystallizes from alcohol in needles, which melt at 218°. The compound is very soluble in hot alcohol, moderately soluble in cold alcohol and insoluble in cold water. Analysis (Kjeldahl):

	Calculated for $C_{12}H_{15}O_3N_2$	Found
N	12.05	12.09

1,3-Dimethyl-4-p-hydroxybenzylhydantoin (*1,3-Dimethyltyrosinehydantoin*),  OH. — Under conditions similar to those employed in the preparation of tyrosine from anisalhydantoin,[1] dimethylanisalhydantoin was digested with hydriodic acid and red phosphorus for nine hours. The excess of hydriodic acid was then evaporated, the residue triturated with water, and the dimethyltyrosinehydantoin then extracted with ether. The hydantoin is very soluble in water and alcohol and moderately soluble in ether. It crystallizes in rhombohedric crystals which melt at 149°–150°. In the preparation of methyltyrosine (below) it is not necessary to isolate this hydantoin. Nitrogen determination (Kjeldahl):

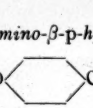
	Calculated for $C_{12}H_{14}O_3N_2$	Found
N	11.96	11.72

1-Methyl-4-p-hydroxybenzylhydantoin (*1-Methyltyrosinehydantoin*),  OH.—Four grams of 1-methyl-4-anisalhydantoin were digested for 2 hours at 150°–160° with 20 cc. of hydriodic acid and one gram of red phosphorus. The excess of hydriodic acid was then evaporated and the residue extracted with dilute alcohol. The solution thus obtained was concentrated to about 10 cc. and then neutralized by addition of ammonia, when the hydantoin separated as an oil, which crystallized on standing. The substance was purified by crystallization from alcohol and separated in short,

[1] Wheeler and Hoffman: *Loc. cit.*

colorless prisms, which melted at 200°. The hydantoin is quite soluble in water. Analysis (Kjeldahl):

	Calculated for $C_{11}H_{12}O_3N_2$	Found
N	12.72	12.37

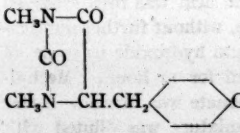
α-Methylamino-β-p-hydroxyphenylpropionic Acid (Methyltyrosine), HO  $CH_2.CH(NHCH_3).COOH$.—Twelve grams of 1,3-dimethyl-4-anisalhydantoin were digested, in an oil bath, with 60 cc. of hydriodic acid and 3.5 grams of red phosphorus. The excess of hydriodic acid was then removed by careful distillation and the residue, without further purification, digested with 70 grams of barium hydroxide in 70 cc. of water. This digestion was continued for 12 hours. Methylamine was evolved and barium carbonate was formed. After the hydrolysis was completed, the mixture was diluted with about 350 cc. of water, heated to boiling and filtered. The excess of barium was then removed from the filtrate by precipitation as carbonate with carbon dioxide and the carbonate removed by filtration. When this solution was concentrated to a small volume and acetic acid was added we obtained 3.5 grams of methyltyrosine, which separated as a brown, crystalline powder. The product was free from barium and decomposed with strong effervescence at 294° on moderately rapid heating. On dissolving the acid again in dilute sodium hydroxide solution and then acidifying with acetic acid the amino acid separated in a pure condition and crystallized in clusters of colorless, prismatic needles. The separation is more complete when an equal volume of alcohol is added to the solution. The acid has no definite melting point. Our product decomposed anywhere between 265° and 320°, according to the rate of heating. Methyltyrosine gives an intense Millon reaction, and also Piria's reaction for tyrosine. Analysis (Kjeldahl):

	Calculated for $C_{10}H_{13}O_3N$	Found
N	7.18	7.05

[1] Ann. Chem. (Liebig), 88, 251.

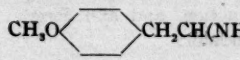
Twelve grams of the 1,3-dimethylanisalhydantoin represents the yield of this compound obtainable from 8.0 grams of hydantoin, therefore the above yield of methyltyrosine corresponds to 22.3 per cent. of the theoretical, calculated from the hydantoin, or to 32.0 per cent. calculated from the amount of anisalhydantoin used. Methyltyrosine is also formed by the action of hydriodic acid on α-methylamino-β-*p*-methoxyphenylpropionic acid (see below). This method of preparation, however, is less suitable for its synthesis than the method described above.

*1,3-Dimethyl-4-*p*-methoxybenzylhydantoin*,



—This hydantoin was obtained, as an oil, by reduction of 1,3-dimethyl-4-anisalhydantoin with tin and hydrochloric acid. The hydantoin was dissolved in hot alcohol, three times the calculated amount of tin added, and hydrochloric acid gas then conducted into the solution until all the tin dissolved. The solution was then evaporated nearly to dryness, triturated with cold water and the aqueous solution extracted with ether. After evaporating the ether we obtained the hydantoin as a light brown oil, which would not solidify. The yield of oil was 61 per cent. of the theoretical. Practically the same yield of the hydantoin was likewise obtained by reduction of the anisalhydantoin with sodium amalgam. The same hydantoin is also obtained by alkylation of 4-anisylhydantoin (see below) with methyl iodide in the presence of alkali. The structure of this hydantoin was established by the compounds formed on hydrolysis.

*α-Methylamino-β-*p*-methoxyphenylpropionic Acid*,



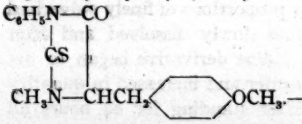
—1,3-Dimethyl-4-anisylhydantoin was hydrolyzed with a hot barium hydroxide solution as described in the preparation of methyltyrosine from its corresponding dimethylhydantoin (see above). Methyl-

amine was evolved during the digestion and barium carbonate was formed. After removing the excess of barium in the usual manner, the solution was then concentrated to a small volume and acetic acid added, when this amino acid immediately deposited. The yields obtained in two experiments were 53 and 57 per cent. of the theoretical, respectively. The acid crystallizes from water in clusters of colorless, prismatic needles, which decompose, with strong effervescence, from 220° to 255°, according to the rate of heating. It is readily soluble in hot water, moderately soluble in cold. Nitrogen determination (Kjeldahl):

	Calculated for $C_{11}H_{15}O_3N$	Found
N	6.71	6.56

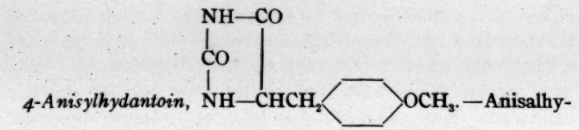
Preparation of Methyltyrosine, $HOC_6H_4CH_2CH(NHCH_3)$-$COOH$, from α-Methylamino-β-p-methoxyphenylpropionic Acid.—Three grams of the amino acid were boiled for 2 hours with 15 cc. of hydriodic acid. Methyl iodide was evolved. The excess of hydriodic acid was then removed by evaporation and the residue triturated with ammonium hydroxide solution to neutralize the hydriodic acid present, when we obtained a yield of methyltyrosine corresponding to 65 per cent. of the theoretical.

3-Methyl-1-phenyl-4-anisyl-2-thiohydantoin,



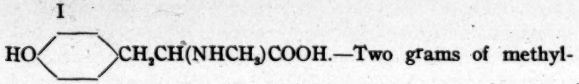
—From phenyl isothiocyanate and α-methylamino-β-*p*-methoxyphenylpropionic acid. The hydantoin crystallizes from dilute alcohol in colorless prisms, which melt at 105° to a clear oil. It is practically insoluble in water, quite soluble in hot and moderately soluble in cold alcohol. It is not soluble in sodium hydroxide solution. Nitrogen determination (Kjeldahl):

	Calculated for $C_{18}H_{18}O_2N_2S$	Found
N	8.59	8.27



4-Anisylhydantoin, —Anisalhydantoin was finely pulverized and suspended in about 40 parts of 95 per cent. alcohol, five equivalents of tin added and hydrochloric acid gas passed into the hot mixture until a clear solution was obtained. The solution was then concentrated by evaporation on the steam bath and diluted with water, when the above hydantoin separated. It was washed with cold water to remove traces of tin chloride and purified. The compound was identical with the hydantoin obtained by reduction of anisalhydantoin with hydriodic acid.[1] The yield was 78 per cent. of the calculated.

α-Methylamino-β-(3,5-diiodo-4-hydroxyphenyl)propionic Acid (Methyliodogorgoic Acid),



—Two grams of methyltyrosine and two molecular proportions of potassium hydroxide were dissolved in 50 cc. of cold water. To the cold solution were then added two molecular proportions of finely pulverized iodine (5.2 grams). The iodine slowly dissolved and after shaking for a few minutes the iodine derivative began to deposit as a brown, crystalline powder and increased in quantity as the iodine disappeared. After standing for 24 hours all the iodine had practically dissolved and we obtained 3.5 grams of the crude iodo acid, which decomposed, on heating rapidly, at about 205° with strong effervescence. The yield was 76 per cent. of the theoretical. The acid is very difficultly soluble in hot water, but dissolves readily in dilute acids and alkalies. It was purified for analysis by dissolving in a one per cent. ammonia solution, decolorizing with bone coal and then reprecipitating by addition of dilute acetic acid. In this

[1] Wheeler, Hoffman and Johnson: *Loc. cit.*

manner the acid was obtained in nearly colorless crystals. The decomposition point is not sharp but varies between 200°-210° according to the rate of heating. Analyses:

	Calculated for $C_{10}H_{11}O_3NI_2$	Found
N	3.13	3.19
I	56.7	55.88

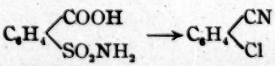
Methyliodogorgoic acid is less soluble in water than iodogorgoic acid and its aqueous solution does not affect litmus or turmeric paper. It dissolves at once in dilute ammonia and also in acids. The acid is not decomposed by cold dilute nitric acid, but concentrated nitric acid liberates iodine at once at ordinary temperature. The acid is precipitated from its dilute sulphuric acid solution by the addition of phosphotungstic acid. On the other hand, it is not precipitated by chloroplatinic acid, picric acid and mercury chloride. It does not give a red color with Millon's reagent.

NEW HAVEN, CONN.
February, 1912

REACTION BETWEEN CARBOXYLIC ACIDS AND BENZENESULPHONAMIDE AT HIGH TEMPERATURES[1]

BY CHAS. A. ROUILLER

In 1889, Remsen and Rich[2] observed that *p*-sulphamidobenzoic acid is converted into *p*-chlorobenzonitrile by the action of phosphorus pentachloride at a high temperature:



This interesting reaction, in which a nitrogen atom breaks loose from a sulphur atom and combines with a carbon atom, was made the object of further study by Hartman,[3] who, by keeping the temperature below 220°, was able to isolate an intermediate product, *p*-cyanobenzenesulphone chloride, $NCC_6H_4SO_2Cl$. Since the sulphamido group does not react readily with phosphorus pentachloride while the carboxamido

[1] Cf. Footnote, THIS JOURNAL, 47, 1.
[2] *Ibid.*, 11, 348.
[3] *Ibid.*, 18, 150; Diss., Johns Hopkins Univ., 1894.

group does, he suggested that the reaction might possibly be due to the isomerization of the sulphamidobenzoic acid into carbamidobenzenesulphonic acid under the influence of the phosphorus pentachloride and heat. On heating *p*-sulphamidobenzoic acid alone for some time near its melting point, he found that it charred considerably and was converted into the acid ammonium salt of *p*-sulphobenzoic acid, while at higher temperatures benzonitrile was obtained.

Muckenfuss[1] confirmed these results, and by regulating the temperature even more carefully was able to isolate still another intermediate product, having the composition *p*-$ClOCC_6H_4SO_2N:PCl_3$, in the reaction between *p*-sulphamidobenzoic acid and phosphorus pentachloride. Moreover, he observed that the acid, which, when heated rapidly, melts at about 290°, fuses if kept a long time at 220°. Under these conditions it does not lose in weight and the resulting product neutralizes practically the same amount of standard alkali as the original acid, but it is entirely different from it in its properties. In fact, Muckenfuss found that it consisted of at least four distinct substances which he was able to isolate in pure condition, namely, free *p*-sulphobenzoic acid, its acid ammonium salt, an acid which he thought was isomeric with *p*-sulphamidobenzoic acid and which he designated as "iso-*p*-sulpaminebenzoic acid," and, finally, a substance having the same composition as the diamide of *p*-sulphobenzoic acid but differing markedly in its properties from this compound. It is neutral, practically insoluble in all solvents except concentrated sulphuric acid and alkalis, and is infusible. The caustic alkalis decompose it easily, liberating all its nitrogen as ammonia and forming salts of *p*-sulphobenzoic acid. Under atmospheric pressure it is unaffected by boiling acids, but in sealed tubes at 200°, hydrochloric acid converts it into acid ammonium *p*-sulphobenzoate and ammonium chloride. The true diamide, on the other hand, prepared from the dipotassium salt of *p*-sulphobenzoic acid through the dichloride, is quite soluble in hot water and in alcohol, melts at 230° and loses only one-half of its nitrogen when boiled with alkalis,

[1] THIS JOURNAL, 18, 150, 349; Diss., Johns Hopkins Univ., 1895.

giving a salt of *p*-sulphamidobenzoic acid. The "iso" acid is very easily soluble. When it is heated with very concentrated sodium hydroxide, no ammonia is evolved until almost all the water has been driven off; on acidifying the solution, sulphur dioxide is set free and ether extracts *p*-hydroxybenzoic acid from the solution. Hydrochloric acid at 270° converts it into acid ammonium *p*-sulphobenzoate.

As found by Hartman and confirmed by Muckenfuss, when *p*-sulphamidobenzoic acid is heated at 285°, there is considerable charring and only acid ammonium *p*-sulphobenzoate, but no "infusible diamide," is formed, whereas at lower temperatures there is practically no charring and the "infusible diamide" appears. In other words, this substance seems to be formed only when the water necessary for the formation of the acid ammonium salt is not supplied by a partial carbonization of the acid. It is not likely that this water can be furnished by such a stable substance as the infusible "diamide," and Muckenfuss thought that it might possibly be furnished by the free *p*-sulphobenzoic acid which must be formed at the same time as the "diamide," and that consequently if sulphobenzoic acid were mixed with the sulphamidobenzoic acid, no "diamide" would be formed on heating the mixture at 220°. Such was found to be the case; only sulphobenzoic acid and its acid ammonium salt were obtained in the product of the reaction.

Stoddard,[1] continuing the work of Muckenfuss, discovered the interesting fact that the "infusible diamide," when boiled with magnesium hydroxide, loses only one-half of its nitrogen as ammonia and yields the magnesium salt of an acid isomeric with *p*-sulphamidobenzoic acid. This acid loses another atom of nitrogen when boiled with a caustic alkali and is probably *p*-carbamidobenzenesulphonic acid. He concluded that both nitrogen atoms of the "infusible diamide" are combined with carbon. He also showed that Muckenfuss' "iso" acid is not isomeric with *p*-sulphamidobenzoic acid but contains one molecule of water less and is, therefore, analogous in compo-

[1] This Journal, **47**, 1.

sition, if not in structure, to saccharin or *o*-benzoic sulphinide. These results were confirmed by Chamberlain.[1]

Waters,[2] working with *p*-sulphamido-*m*-toluic acid, and Nakaseko,[3] using *m*-sulphamidobenzoic acid, found that these acids, when heated, undergo changes similar to those observed in the case of *p*-sulphamidobenzoic acid, except that they obtained no evidence of the formation of products analogous to Muckenfuss' "iso" acid.

Nakaseko, whose work was done in this laboratory in 1899, was the first to suggest that the "infusible diamide" might be an amidine compound, $HO_3SC_6H_4C(:NH)NH_2$. Such a constitution explains all the known reactions of the substance, its neutral reaction, slight solubility and infusibility probably being due to the formation of an inner salt. Simmons[4] adopted the same view.

In order to test the correctness of Nakaseko's suggestion, Robinson[5] attempted to prepare the substance from potassium *p*-cyanobenzenesulphonate by the Pinner and Klein method for the preparation of amidines from nitriles, but was unable to bring about any reaction between the cyano salt, absolute alcohol and gaseous hydrochloric acid, other than the liberation of cyanobenzoic acid and the formation of potassium chloride. He then treated benzamidine hydrochloride in alcohol with silver benzenesulphonate, and found the resulting benzamidine benzenesulphonate, $C_6H_5C(:NH)NH_2.HO_3SC_6H_5$, to be a well defined, crystalline substance, fairly soluble in cold water and melting at 174°. He did not pursue the investigation further.

At this point, Professor Remsen suggested to the writer that, instead of heating a substance in which both the sulphamido and the carboxyl groups are present in the same molecule, he try heating a mixture of a carboxylic acid, like benzoic acid, with a sulphonamide, such as benzenesulphonamide. When this was done, the main product of the reaction was

[1] THIS JOURNAL, **47**, 318.
[2] *Ibid.*, **47**, 333.
[3] *Ibid.*, **47**, 429.
[4] Diss., Johns Hopkins Univ., 1901.
[5] *Ibid.*, 1906, p. 74.

found to be identical with Robinson's benzamidine benzenesulphonate. This seemed to afford a new general method for the preparation of amidines, and in fact a number of amidine benzenesulphonates derived from both aliphatic and aromatic acids have been prepared in this way, but, as will be shown in the experimental part, in a number of cases negative results were obtained. The behavior of the three phthalic acids when heated with benzenesulphonamide is interesting. *o*-Phthalic acid gives phthalimide almost quantitatively, while the chief products obtained from the *m*- and *p*-acids were the corresponding dinitriles, together with very small quantities of the cyanobenzoic acids. The *p*-acid also yielded some *p*-cyanobenzamidine benzenesulphonate, $NCC_6H_4C(:NH)NH_2.HO_3SC_6H_5$.

It now became of interest to learn whether the infusible diamide could be obtained by heating *p*-sulphobenzoic acid with benzenesulphonamide. Such was found to be the case. This seems to show pretty conclusively that the substance really has the structure suggested by Nakaseko. That carbamidobenzenesulphonic acid is an intermediate product in its formation is indicated by the fact that benzamidine benzenesulphonate can be obtained from benzamide and benzenesulphonamide, but the presence of a free acid (benzenesulphonic acid was used in this case) seems to be essential.

The true constitution of Muckenfuss' "iso" acid is still doubtful. The careful analyses of its salts by Stoddard and Chamberlain seem to show beyond doubt that it is isomeric with saccharin, but its structure as the *p*-analogue of saccharin cannot be accepted without further evidence. All attempts by the writer to alkylate it by means of dimethyl sulphate were unsuccessful.

Experimental Part

Preliminary Experiments with Benzoic Acid and Benzenesulphonamide.—Equimolecular quantities of the acid and amide were intimately mixed in a mortar, and small portions of the mixture heated in sealed tubes 5 to 6 hours at different temperatures between 200° and 300°. The contents of the tubes were then distilled, first with 4 cc. of 2.5 N caustic soda

and 4 cc. of 5 N magnesium chloride, then with 10 grams of caustic soda. The amount of ammonia set free in each operation was titrated with cochineal as the indicator. As the sulphonamide group is not attacked by either magnesium or sodium hydroxide, the carboxamide group only by sodium hydroxide, while the infusible compound formed when *p*-sulphamidobenzoic acid is heated loses half of its nitrogen when boiled with magnesium hydroxide and the other half when distilled with caustic soda, it was hoped that the above treatment of the reaction product might furnish some indication as to the extent of any reaction which might have taken place and as to the most favorable conditions of time, temperature, etc. The small tubes were heated in beakers containing sulphuric acid and acid potassium sulphate, and it was found impossible to maintain the temperature perfectly constant during the 5- to 6-hour periods of heating. Moreover, as was found afterwards, benzamidine benzenesulphonate is one of the products of the reaction and the mixture of magnesium chloride and caustic soda slowly liberates more than one-half of the nitrogen of this substance. It is not surprising, therefore, that there should be some exceptions to the general trend of the results and that the duplicate determinations should vary considerably. The results (expressed in percentages of the total nitrogen present in the mixture) are tabulated below. Those which are obviously erroneous are enclosed in parentheses.

Temp.	Length of heating Hours	Length of boiling with $Mg(OH)_2$ Hours	Nitrogen set free $Mg(OH)_2$	Nitrogen set free NaOH
200°	6	1	27.6	1.8
	5.75	1.25	(40.9)	(6.7)
	5.5	2.25	22.7	1.1
225°	6	1	36.5	38.0
	5.75	1.25	35.2	(53.1)
	5.5	2.25	41.5	42.5
250°	6	1	38.2	38.1
	5.75	1.25	43.7	34.0
275°	6	1	(77.8)	24.6
	5.75	1.25	55.5	21.9
300°	6	1	60.8	22.9
	5.25	1.25	59.4	17.3

The mixtures melted at about 115°–120° to a clear liquid, and below 200° there was apparently no reaction. Above that temperature the liquid gradually darkened and at 300° was a very dark brown. On cooling, the product became lighter in color and was of the consistency of hard butter. On opening the tubes there was found to be some slight pressure (one tube heated above 300° exploded) and the odor of benzonitrile was plainly perceptible. At the higher temperatures this compound was formed in such quanatities that, on cooling, a layer of it separated above the solidified reaction product. From the table above it will be seen that after heating at 200° only about 27 per cent. of the nitrogen is set free by distilling successively with magnesium and sodium hydroxides; at 225° and 250° about 79 per cent.; at 275°, 76 per cent.; and at 300°, 80 per cent.; i. e., there appears to be no further appreciable change of this nature above 225°. The amount of nitrogen liberated by the magnesium hydroxide, however, increases gradually from about 25 per cent. for the mixture heated at 200° to 60 per cent. for that heated at 300°, while the nitrogen set free by caustic soda jumps suddenly from about 1.5 per cent. at 200° to 40 per cent. at 225°, then gradually falls to 20 per cent. at 300°. The temperature 225° was, therefore, chosen as being most promising for subsequent experiments. It happens, too, that this is the most favorable temperature for the formation of the infusible product from *p*-sulphamidobenzoic acid, and subsequent work has shown that this seems to be generally true for all cases.

Reaction at 225°.—Larger quantities of the mixture were accordingly heated at 225° for several hours. The reaction product was then distilled with steam until the distillate no longer smelled of benzonitrile. The distillate was neutralized with caustic potash, repeatedly extracted with ether, the ethereal extract evaporated and the benzonitrile weighed. The amount of benzonitrile obtained in various experiments from equimolecular mixtures of benzoic acid and benzenesulphonamide heated 6 hours at 225° was 3.0–3.7 per cent. of the weight of the mixture.

The alkaline distillate, after removal of the nitrile, was

concentrated and acidified and the precipitated benzoic acid combined with that obtained from the solution remaining in the distilling flask.

After removal of the benzonitrile and part of the benzoic acid by distillation with steam as described above, there remained in the flask a small quantity of a very difficultly soluble substance which was filtered off while the liquid was still boiling hot. The yield of this compound, which was identified as kyaphenin, $(C_6H_5CN)_3$, was 2.3–2.9 per cent. of the weight of the original mixture of benzoic acid and benzenesulphonamide.

The filtrate from the kyaphenin, on cooling, deposited considerable benzoic acid, and a further small quantity was obtained by extracting the cold solution with ether. It was combined with that obtained from the steam distillate and amounted to 42–52 per cent. of that originally present in the mixture.

The solution, after the removal of the benzoic acid, was still strongly acid. It was neutralized with barium hydroxide and the small amount of barium sulphate (0.2–0.4 gram from 50 grams of the original mixture) was filtered off.

The filtrate contained a large quantity of a barium salt and gave off ammonia copiously when boiled with alkalis. On concentation, pearly leaflets of barium benzenesulphonate were deposited.

On further concentration, needles resembling synthetic benzamidine benzenesulphonate (see page 483) and melting at about 170° were obtained. They gave off ammonia with alkalis but still contained barium, and all attempts to purify them completely by fractional crystallization from water were unsuccessful. In all subsequent experiments, therefore, the filtrate from the barium sulphate was evaporated to dryness and the residue extracted with absolute alcohol. The portion insoluble in alcohol was found to be practically pure barium benzenesulphonate. From 43 grams of the original mixture, 15.6 grams of this salt were obtained.

The alcoholic extract was evaporated to dryness. The residue was almost pure benzamidine benzenesulphonate.

It melted at 170°–180° and after recrystallization from water was obtained in radiating needles melting at 173°. A mixture with the synthetic product (melting at 172°–174°) melted at 172°–174°.

NH_3 liberated from 0.3293 gram substance by boiling NaOH neutralized 24.70 cc. of 0.0956 N H_2SO_4.

0.1935 gram substance gave 0.1597 gram $BaSO_4$ (Carius).

	Calc. for $C_6H_5C(:NH)NH_2.HO_3SC_6H_5$	Found
N	10.07	10.05
S	11.53	11.34

As shown above (page 482), about one-half of the benzoic acid was recovered unchanged. An experiment was, therefore, carried out with a mixture of one molecule benzoic acid and *two* molecules benzenesulphonamide. The reaction product was treated in exactly the same way as described above. The results may be summarized as follows:

8.9 grams of the mixture gave
- 0.37 gram benzonitrile
- 0.04 gram kyaphenin
- 0.09 gram benzoic acid
- trace of sulphuric acid
- 2.79 grams benzenesulphonic acid
- 4.35 grams benzamidine benzenesulphonate.

The main reaction which had taken place can therefore be represented by the following equation:

$$C_6H_5COOH + \begin{matrix} H_2NO_2SC_6H_5 \\ H_2NO_2SC_6H_5 \end{matrix} = C_6H_5C\begin{matrix} \diagup\!\!\diagup NH \\ \diagdown NH_2.HO_3SC_6H_5 \end{matrix} + HO_3SC_6H_5$$

Calc.	2.5 grams	6.4 grams	5.7 grams	3.2 grams
Found			4.3 grams	2.8 grams

Benzamidine Benzenesulphonate.—This substance was made by Robinson[1] from benzamidine hydrochloride and silver benzenesulphonate. It crystallizes from water in radiating clusters of needles or from very concentrated solutions in rosettes of minute crystals melting at 173°. It is very soluble in water and alcohol and is precipitated from the latter solvent by ether. When boiled with caustic alkalis, all the nitrogen is set free as ammonia. Boiling with a suspension of solid magnesium hydroxide, *free from alkali salts*, in water liberates

[1] Diss., Johns Hopkins Univ., 1906, p. 77.

only half of the nitrogen, as shown by the following experiment:

0.2148 gram was boiled with about twice the calculated amount of $Mg(OH)_2$ and the ammonia set free titrated with 0.0956 N H_2SO_4. Calc. for one-half of the N, 8.07 cc. acid.

Length of boiling	Acid used	Total acid used
1st hour	7.61	7.61
2nd hour	0.49	8.10
3rd hour	0.18	8.28
4th hour	0.19	8.37
5th hour	0.06	8.43

That is, 50.2 per cent. of the total nitrogen was set free in the first two hours of boiling and only 2.0 per cent., due to the slow decomposition of the benzamide which is formed, is liberated in the next three hours. In another experiment, the ammonia from 0.2968 gram of the salt boiled with excess of magnesium hydroxide for a few hours neutralized 10.23 cc. instead of the calculated 10.95 cc. of acid, and the solution, after being filtered and concentrated, deposited crystals of benzamide melting at 124°–126°.

Later, Dr. E. Emmett Reid, working in this laboratory, found that benzamidine hydrochloride, when boiled with his mixture of magnesium chloride and sodium hydroxide,[1] loses almost all of its nitrogen as ammonia, the rate at which the ammonia is evolved decreasing with the length of boiling. Repeating the experiment under the same conditions with benzamidine benzenesulphonate, I obtained similar results, as shown in the following table:

0.1641 gram benzamidine benzenesulphonate was boiled with 4 cc. of 5 N $MgCl_2$ and 4 cc. 2.5 N NaOH and the NH_3 titrated at the end of each hour with 0.0956 N H_2SO_4. Calc. for half of the N, 6.15 cc. acid.

Length of boiling	Acid used	Total acid used
1st hour	3.28	3.28
2nd hour	3.05	6.33
3rd hour	1.41	7.74
4th hour	0.61	8.35
5th hour	0.32	8.67
6th hour	(0.76)	(9.43)
7th hour	0.28	(9.71)

[1] THIS JOURNAL, **21**, 293.

Thus, although approximately half of the nitrogen is set free in the first two hours, considerably more than half has been liberated at the end of the third hour and even during the seventh hour an appreciable amount of ammonia is evolved. These facts furnish an interesting illustration of the influence of neutral salts on reactions and deserve investigation.

As the process described above affords an easy method for the preparation of a benzamidine salt, attempts were made to obtain the hydrochloride from the benzenesulphonate by adding the calculated amount of hydrochloric acid to an aqueous solution of the latter and concentrating, but only the unchanged sulphonate was obtained. The solution was then neutralized with barium hydroxide, but benzamidine benzenesulphonate was again obtained on concentrating.

Benzamidine benzenesulphonate dissolves easily in concentrated sulphuric acid with a yellow color which disappears when the solution is diluted with water. The solution in concentrated acid was heated to 200°, cooled, neutralized with barium hydroxide, filtered and evaporated to dryness. The residue contained a small amount of barium salt insoluble in absolute alcohol (barium benzenesulphonate), while ether precipitated unchanged benzamidine benzenesulphonate (melting at 175°) from the alcoholic extract. When the concentrated sulphuric acid solution was kept at 200° for one hour, a sublimate of benzoic acid was deposited on the upper part of the vessel and no benzamidine salt was recovered.

Boiled with 10 cc. dilute hydrochloric acid under a return condenser for 6.5 hours, 0.66 gram benzamidine benzenesulphonate lost 0.01 gram in weight. The dry product melted at 165° and was soluble in absolute alcohol, the solution depositing the unchanged amidine salt melting at 173°–175°. Similar results were obtained when the salt was evaporated to dryness twice with concentrated hydrochloric acid. There was no change in weight and the residue melted at 170°, and after one crystallization from alcohol at 175°.

The behavior of the salt towards dilute and concentrated nitric acids is entirely analogous. Although a slight odor of nitrobenzene was perceived during the evaporation to dry-

ness, there was no appreciable change in weight, and the residue, which melted at 165°, dissolved easily in alcohol with a yellow color and the unchanged amidine salt was recovered from the solution.

Heated in a sealed tube with dilute hydrochloric acid at about 200° for 5.5 hours, the salt gave benzoic acid, ammonium chloride and benzenesulphonic acid. Two other substances were also formed, but in too small quantity for identification. One gives with barium salts in alkaline solution a substance which is insoluble in water but easily soluble in acids and gives no precipitate with sulphates in neutral solution. The second is soluble in alcohol, gives off ammonia with alkalis and although it softens and contracts at about 210°, it is infusible below 300°.

Benazmidine benzenesulphonate is thus, in its behavior towards caustic alkalis, magnesium hydroxide and acids, entirely analogous to the infusible product obtained by heating *p*-sulphamidobenzoic acid.

Formation of Benzamidine Benzenesulphonate from Benzamide and Benzenesulphonamide.—The formation of the amidine salt from benzoic acid and benzenesulphonamide may be represented as taking place in two stages:

$$C_6H_5COOH + H_2NO_2SC_6H_5 = C_6H_5CONH_2 + HO_2SC_6H_5$$
$$C_6H_5CONH_2 + H_2NO_2SC_6H_5 = C_6H_5C(:NH)NH_2.HO_2SC_6H_5$$

In order to determine whether benzamide is really an intermediate product in the reaction, equimolecular proportions of benzamide and benzenesulphonamide were heated together at 220°. The product was practically colorless, although the odor of benzonitrile was perceptible, and on treatment with boiling magnesium hydroxide only 0.7 per cent. of the nitrogen was set free as ammonia. Alongside of the tube containing the mixture of amides was heated another containing the same amounts of amides and in addition about one molecule of free benzenesulphonic acid which, although it had been heated at 110° and allowed to cool over sulphuric acid, contained nearly a molecule of water. The contents of this tube, after being heated, were black and smelled strongly of benzonitrile. Magnesium hydroxide liberated 51.8 per cent.

of the nitrogen present in the original mixture and caustic soda 19.3 per cent. In another experiment with somewhat larger quantities of material, the reaction product was poured into water, extracted with ether to remove the benzonitrile and the water solution concentrated to a small volume. Crystals were deposited which, when dissolved in alcohol and reprecipitated with ether, melted at 172–174°, and the melting point was not altered by mixing the crystals with benzamidine benzenesulphonate obtained in a different manner. It seems, therefore, that free benzenesulphonic acid must be present before the reaction can take place. That the water present in the acid had little, if any, influence was shown by heating a mixture of the amides (one molecule of each) with one molecule of water; there was apparently no reaction and magnesium hydroxide liberated only 2.6 per cent. of the nitrogen; the slight increase in this value as compared with that for the mixture containing no water, 0.7 per cent. (see above), can be easily explained by the hydrolyzing action of the water on the amides in the sealed tube at 220°.

Benzamidine p-*Toluenesulphonate.*—Robinson[1] states that when benzamidine hydrochloride and silver *p*-toluenesulphonate are brought together in alcoholic solution and the silver chloride is filtered off, the solution deposits colorless, flattened prisms having a very high melting point (above 310°). This supposed benzamidine *p*-toluenesulphonate was not studied further.

This further similarity of high melting point to the infusible product obtained from *p*-sulphamidobenzoic acid made it desirable to study this compound somewhat more fully. Equivalent quantities of benzamidine hydrochloride and silver *p*-toluenesulphonate in aqueous solutions were mixed together and the silver chloride filtered off. On concentrating the solution, the benzamidine *p*-toluenesulphonate separated out as a coarse crystalline powder or, from very concentrated solutions, in spherical clusters of minute crystals which, on standing, became quite hard. To my surprise, however, the salt melted sharply at 193°.

[1] Diss., Johns Hopkins Univ., 1906, p. 77.

A preliminary experiment having shown that when a mixture of benzoic acid with 2 molecules of *p*-toluenesulphonamide is heated 5.25 hours at 220°, magnesium and sodium hydroxides liberate 58.8 and 22.2 per cent., respectively, of the original nitrogen from the reaction product, the experiment was repeated on a somewhat larger scale and the product treated in the manner described above for the mixtures of benzoic acid and benzenesulphonamide. The amidine salt so obtained melted at 190° (owing to the small amount available, it could not be completely purified), and a mixture with the synthetic product melted at 190°-191°. Analyses of synthetic product:

NH_3 from 0.2230 gram salt boiled with NaOH neutralized 15.54 cc. of 0.1017 N H_2SO_4.
NH_3 from 0.1823 gram salt boiled with $Mg(OH)_2$ and NaOH neutralized 6.02 and 6.40 cc. 0.1011 N H_2SO_4, respectively.

	Calc. for $C_6H_5C(:NH)NH_2.HO_3SC_6H_4CH_3$	Found	
		$Mg(OH)_2$	NaOH
Total N	9.59		9.93
Half N	4.79	4.72	4.86

Acetamidine Benzenesulphonate.—When the usual mixtures (one molecule acetic acid and two of benzenesulphonamide) are heated, they melt at about 150° but below 200° there is apparently no reaction; above 200° the liquid begins to darken, and when kept at 215°–225° for some hours, a black tarry product is formed.

Temp.	Time Hours	Nitrogen set free by $Mg(OH)_2$	NaOH
210°	3	4.1	0.6
195°	2	48.4	15.9
215°	3.5		
220°	5	57.9	34.4

In these experiments the acetic acid was made up to 100 per cent. by titrating the laboratory acid with a standard alkali and adding the calculated amount of acetic anhydride. Having found from the preliminary experiments tabulated above that about 220° seemed to be the most favorable tempera-

ture for the reaction, the heating at this temperature was repeated with somewhat larger quantities of materials. The product was treated in the usual way and finally a substance was obtained which was extremely easily soluble in water and alcohol, crystallizing from concentrated water solution in long, transparent needles which melted at 136°. Analysis:

NH_3 set free from 0.1513 gram substance by $Mg(OH)_2$ and NaOH neutralized 6.76 and 6.75 cc., respectively, of 0.1021 N H_2SO_4.

	Calc. for $CH_3C(:NH)NH_2.HO_3SC_6H_5$	Found	
		$Mg(OH)_2$	NaOH
Half N	6.48	6.39	6.38

Benzenesulphonamide and Oxalic Acid.—Mixtures of 1 molecule of the acid and four of the amide were heated in the usual way. They melted at about 150° with evolution of gas. As the temperature was raised the product darkened only slightly. There was considerable pressure in the tubes.

Temp.	Time Hours	Nitrogen set free by $Mg(OH)_2$	NaOH
150°–180°	2	2.4	0.1
200°	2		
215°	5.5	10.5	0.0
250°	5.5	21.6	1.0

The above table gives no evidence of amidine formation, as the amount of ammonia liberated by caustic soda is practically *nil*.

An attempt was now made to determine whether the method of preparing amidines which has been described above is at all general, and with this end in view benzenesulphonamide was heated with such organic acids as were on hand, the products distilled two hours with magnesium hydroxide and then one hour with excess of caustic soda. The results, expressed in percentages of the nitrogen liberated by each reagent, are tabulated below.

Acid	Nitrogen set free by $Mg(OH)_2$	NaOH
*Malonic	54.1	15.1
Tartaric	20.0	1.9
Chloroacetic	67.8	1.6
o-Phthalic	55.3	24.1
**m*-Phthalic	49.0	36.1
**p*-Phthalic	45.7	28.0
Salicylic	5.1	0.6
o-Methoxybenzoic	39.0	5.5
p-Hydroxybenzoic	17.2	2.0
o-Nitrobenzoic	61.3	3.5
**m*-Nitrobenzoic	44.0	25.7
**p*-Nitrobenzoic	49.4	21.8
o-Aminobenzoic	9.3	0.4
p-Aminobenzoic	12.9	3.0
**p*-Sulphobenzoic	72.3	21.7
*Phenylacetic	66.1	23.5
3-Nitrophthalic	55.7	5.3
**m*-Bromobenzoic	47.4	9.0

In those cases (marked with an asterisk in the table above) where the amount of nitrogen liberated by caustic soda seemed to indicate appreciable amidine formation, the experiments were repeated on a larger scale, and in every case except those of malonic and the phthalic acids the corresponding amidines were isolated. The reaction, though applicable to a number of acids, does not, however, seem to be general. Whether steric hindrance plays any part in the reaction, as might be concluded from the fact that *m*- and *p*-nitrobenzoic acids easily give the corresponding amidines, while the ortho acid does not, cannot as yet be determined from the data at hand. This point is now being investigated. As shown in the preceding articles, *p*-sulphamidobenzoic, *m*-sulphamidobenzoic and *p*-sulphamido-*m*-tolucic acids, in which the ortho positions to the carboxyl group are not occupied, all form an amidine (inner?) salt when heated. Mr. Nowell and the writer are now engaged in applying the same reaction to *p*-sulphamido-*o*-toluic and 2,6-dinitro-4-sulphamidobenzoic acids, in which one and both ortho positions, respectively, are occupied. In this connection it is interesting to recall that *o*-substituted amidines cannot be obtained by the Pinner method.

Malonic Acid and Benzenesulphonamide.—Twenty millimoles of the acid (2.08 grams) and 80 of the amide were heated 5 hours at 220°. There was considerable pressure in the tube. The reaction product was black and almost entirely soluble in hot water. The solution was neutralized with barium hydroxide and evaporated to dryness, and the residue extracted with absolute alcohol. The alcoholic extract deposited some benzenesulphonamide and finally, when concentrated to a very small volume, ill-defined crystals, melting at about 195°–205°, which gave off ammonia when boiled with magnesium or sodium hydroxides and were precipitated from alcohol by ether, but the small amount obtained and the exhaustion of the supply of malonic acid at hand made it impossible to determine whether the substance was the malonamidine benzenesulphonate sought.

m-*Nitrobenzamidine Benzenesulphonate.*—This salt, obtained by heating *m*-nitrobenzoic acid with benzenesulphonamide and concentrating the water solution of the reaction product, forms plates melting at 198°–200°. Yield, 1.2 grams from 3.4 grams of acid. Analysis:

NH_3 set free from 0.2068 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 6.10 and 6.05 cc. of 0.1021 N H_2SO_4.

	Calc. for $O_2NC_6H_4C(:NH)NH_2.HO_3SC_6H_5$	Nitrogen set free by $Mg(OH)_2$	NaOH
1/3 N	4.33	4.21	4.18

p-*Nitrobenzamidine Benzenesulphonate.*—From 3.34 grams of *p*-nitrobenzoic acid and 6.28 grams of benzenesulphonamide were obtained 0.8 gram of *p*-nitrobenzonitrile and 2.0 grams of the amidine salt which crystallizes from water in minute needles grouped into spherical aggregates and from alcohol containing a little water in pearly scales. It melts at 250°. Analysis:

NH_3 liberated from 0.2196 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 6.75 and 6.58 cc. 0.1021 N H_2SO_4.

	Calc. for $O_2NC_6H_4C(:NH)NH_2.HO_3SC_6H_5$	Nitrogen set free by $Mg(OH)_2$	NaOH
1/3 N	4.33	4.40	4.28

m-*Bromobenzamidine Benzenesulphonate.*—One and three-tenths grams of this salt were obtained from 4 grams of *m*-bromobenzoic acid. It crystallizes from concentrated water solution in slightly yellow spherical aggregates and melts at 156°–158°. Analysis:

NH_3 set free from 0.2635 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 7.32 and 7.30 cc. 0.1021 N H_2SO_4.

	Calc. for $BrC_6H_4C(:NH)NH_2.HO_3SC_6H_5$	Nitrogen set free by $Mg(OH)_2$	Nitrogen set free by NaOH
1/2 N	3.92	3.97	3.96

Phenylacetamidine Benzenesulphonate.—This salt separates from concentrated water solutions in minute crystals melting at 182°–183°. Yield, 2.5 grams from 2.72 grams of phenylacetic acid. Analysis:

NH_3 liberated from 0.2609 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 8.77 and 8.68 cc. 0.1021 N H_2SO_4.

	Calc. for $C_6H_5CH_2C(:NH)NH_2.HO_3SC_6H_5$	Nitrogen set free by $Mg(OH)_2$	Nitrogen set free by NaOH
1/2 N	4.79	4.70	4.65

p-*Sulphobenzoic Acid and Benzenesulphonamide.*—It seemed probable that in the reaction between these two substances the amidine group, if formed, would combine with the sulphonic acid group in the same molecule and not with that resulting from the benzenesulphonamide; in other words, that the product would be identical with that obtained by heating *p*-sulphamidobenzoic acid. Experiment confirmed this. Five and five-hundredths grams of the acid were heated with 7.85 grams of the amide (2 molecules) and the reaction product treated with water. There were thus obtained 1.2 grams of a product very difficultly soluble in most solvents, but easily soluble in alkalis, from which acids reprecipitated it in crystalline form. It did not melt below 300° and was in all respects identical with the product obtained from *p*-sulphamidobenzoic acid. Analysis:

NH_3 set free from 0.2008 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 10.04 and 9.79 cc. of 0.1021 N H_2SO_4.

	Calc. for C_6H_4<$C(:NH)NH_2$ / SO_3H	Nitrogen set free by $Mg(OH)_2$	Nitrogen set free by NaOH
1/2 N	7.00	7.15	6.97

o-*Phthalic Acid and Benzenesulphonamide.*—Twenty millimoles (3.32 grams) of the acid were heated with 80 millimoles of the amide. The product was practically completely soluble in boiling water, but on cooling 2.72 grams of a substance separated out in fine needles melting at 229°. The characteristic appearance of the needles, the melting point and the fact that this was not changed by mixing the substance with imide from another source, all showed that this substance was phthalimide. The yield was almost quantitative (93 per cent.).

One of the products obtained when *p*-sulphamidobenzoic acid is heated has the empirical composition of an ammonium sulphinide, $C_6H_4\langle^{CO}_{SO_2}\rangle N.NH_4$ (see preceding papers), but as no other compounds are known in which two *para* substituents are thus joined by a nitrogen (or oxygen) atom, and as *o*-phthalimide is so easily obtained from *o*-phthalic acid and benzenesulphonamide, it appeared of interest to try the same reaction with *m*- and *p*-phthalic acids. There is no apparent reason why, if the compound $p\text{-}C_6H_4\langle^{CO}_{SO_2}\rangle NH$ is capable of existence, the analogous compounds, *m*- and $p\text{-}C_6H_4\langle^{CO}_{CO}\rangle NH$ could not also be made. The hope of being able to prepare such compounds was not realized, but rather interesting results were nevertheless obtained in the attempt.

m-*Phthalic Acid and Benzenesulphonamide.*—Three and thirty-two hundredths grams of the acid and 12.56 grams of the amide (4 molecules) were intimately mixed and heated at 223°–238° until the mixture was homogeneously liquid (30 minutes). After the contents of the tube had cooled, they were ground up in a mortar with water several times and filtered. The insoluble portion (1.05 grams) was suspended in water, just neutralized with standard alkali and again filtered. There remained on the filter 0.9 gram of a substance melting at 150°–153°, which was completely soluble in alcohol and was identified as *m*-phthalic dinitrile. The original

water filtrate, on extraction with ether, yielded 0.3 gram more of the dinitrile and 0.2 gram of an acid soluble in ether but precipitated by hydrochloric acid from a small volume of water. It melted at 200°–205°. Not enough was obtained for a complete analysis, but a titration with standard alkali gave a molecular weight of 147, while that of *m*-cyanobenzoic acid, which is described in the literature as melting at 217°, is 152. While concentrating the original water filtrate a cresollike odor was noticed and the vapors arising from the liquid were found to be distinctly acid to litmus. The solution was accordingly distilled with steam and the distillate collected in a vessel containing caustic soda. A sodium salt was thus obtained, but in such small quantity that it could not be identified. The liquid remaining in the flask was neutralized with barium hydroxide and the solution evaporated to dryness; the residue weighed 15.4 grams and consisted of barium and ammonium benzenesulphonates. Absolute alcohol dissolved about 8 grams of it, the solution depositing the ammonium salt in crystals melting at 272°–275°. This melting point was not altered by mixing the substance with ammonium benzenesulphonate prepared in another way. The residue insoluble in alcohol, when treated in water with ammonium sulphate, yielded this same ammonium salt.[1]

p-*Phthalic Acid and Benzenesulphonamide.*—*p*-Phthalic acid gave results analogous to those given by the meta compound. Thus, from 5 grams of the acid were obtained 1.85 grams of *p*-phthalic dinitrile, melting at 221°, a very small amount (0.15 gram) of an acid soluble in ether, having a very high melting point (above 300°) and an apparent molecular weight of 163 instead of 152 calculated for *p*-cyanobenzoic acid, and 22.2 grams of residue (after neutralization of the solution

[1] Morton (J. Am. Chem. Soc., **19**, 837) states that ammonium benzenesulphonate begins to decompose at 236° and melts at 256°. The salt was prepared by me in three ways—by boiling benzenesulphone chloride with water, evaporating to dryness several times, adding ammonia and crystallizing from water, alcohol or alcohol-ether; from benzenesulphonic acid and ammonia; and from barium benzenesulphonate and ammonium sulphate. In every case, the product obtained began to soften at 260° but did not melt below 271°–275°. The values obtained for the percentages of nitrogen in the various preparations (determined by distilling the salt with magnesium or sodium hydroxides and titrating the ammonia) were somewhat low, viz., 7.50, 7.46, 7.72, 7.68; the calculated value is 7.98.

with barium hydroxide and evaporation to dryness). From this residue were obtained, as in the case of the meta acid, barium and ammonium benzenesulphonates. It was noticed, however, that the first fractions of crystals deposited from the absolute alcohol extract of this residue melted *lower* than the succeeding fractions, and repeated recrystallizations further lowered the melting point until a fraction with a constant melting point of 215°–218° was obtained. It was suspected that this product might be the amidine salt corresponding to the cyanobenzoic acid found in small quantity among the products of the reaction, viz., p-*cyanobenzamidine benzenesulphonate*, $NCC_6H_4C(:NH)NH_2.HO_3SC_6H_5$. A compound of this structure should lose one-third of its nitrogen as ammonia when boiled with magnesium hydroxide and two-thirds with caustic soda. The results of analyses agreed equally well for an amide amidine salt, $H_2NOCC_6H_4C(:NH)NH_2.HO_3SC_6H_5$, as for the cyano compound:

I. NH_3 set free from 0.3853 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 12.03 and 23.67 cc. of 0.1011 N H_2SO_4.

II. NH_3 set free from 0.2686 gram salt by $Mg(OH)_2$ and NaOH, respectively, neutralized 8.92 and 16.72 cc. of 0.1011 N H_2SO_4.

	Calc. for $NCC_6H_4C(:NH)NH_2.HO_3SC_6H_5$	Calc. for $H_2NOCC_6H_4C(:NH)NH_2.HO_3SC_6H_5$	
1/3 N	4.62	4.36	
2/3 N	9.24	8.72	
		Found	
		I	II
With $Mg(OH)_2$		4.42	4.70
With NaOH		8.70	8.81

That the substance was really the cyano compound was determined in the following manner: Two grams of it were boiled with magnesium hydroxide until no more ammonia was evolved. If the substance were an amide amidine, it should have given *p*-phthalic diamide; if it were a cyano amidine it should form p-*cyanobenzamide*. Analysis showed that the latter was the compound formed. The product separated in fine needles having exactly the same appearance and melting

point (223°) as *p*-phthalic dinitrile, but when mixed with the latter substance the melting point was lowered 20°–35°. Analysis:

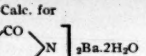
NH_3 set free from 0.1912 gram substance by NaOH neutralized 25.57 cc. of 0.1011 N H_2SO_4.

	$C_6H_4(CONH_2)_2$	Calc. for $NCC_6H_4COHN_2$	$C_6H_4(CN)_2$	Found
N	16.92	19.18	21.90	18.94

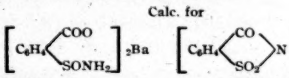
Attempts to Methylate the Supposed p-*Benzoic Sulphinide.*—The experiments with the phthalic acids having failed to give a *m*- or *p*-pthalimide analogous to the supposed sulphinide obtained by heating *p*-sulphamidobenzoic acid, as described in the preceding papers, an attempt was made to methylate this sulphinide in the hope of obtaining analogues of the alkylsaccharins or *o*-sulphinides. *p*-Sulphamidobenzoic acid was heated in test tubes (which were stoppered with corks after their contents had reached the temperature of the bath) in an oil bath at 225° for 8 hours and the product extracted with cold water by grinding up in a mortar several times and filtering until the filtrate was neutral. The solution was then neutralized with barium hydroxide and the difficultly soluble barium sulphinide repeatedly crystallized from large quantities of boiling water. Although the salt is very difficultly soluble and separates in beautiful, apparently pure, silky needles, it does not give concordant results on analysis. It holds water with great tenacity, slowly losing weight even after being heated several hours as high as 230°. The loss in weight varies for different samples; that it is due to loss of water of crystallization or occluded water, not to a change in structure, seems probable from the fact that the heating does not change the properties of the salt. Analyses:

I. 0.1019 gram of a sample which had been heated several hours at about 150° gave 0.0457 gram $BaSO_4$.
II. 0.2806 gram air-dried salt lost 0.0278 gram after several hours at about 150° and then gave 0.1116 gram $BaSO_4$.
III. 0.3539 gram air-dried salt lost 0.0261 gram after several hours at 120° and then gave 0.1436 gram $BaSO_4$.

IV. 0.5062 gram air-dried salt lost 0.0561 gram after 1.5 hours at 130°, 1.5 hours at 200° and 2 hours at 230° and then gave 0.2023 gram $BaSO_4$.



	Calc. for	Found		
		II	III	IV
$2\,H_2O$	15.29	9.91	7.37	11.09



	Calc. for		Found			
			I	II	III	IV
Ba	25.55	27.38	26.41	25.98	25.78	26.44

This barium salt was suspended in methyl alcohol, and dimethyl sulphate was added. The mixture was shaken vigorously at frequent intervals and then allowed to stand overnight and finally heated on the water bath. The aqueous solution thus obtained, when filtered from the barium sulphate, was slightly acid and on the addition of barium hydroxide, the characteristic needles of the barium sulphinide were obtained.

The same results were obtained when a solution of the sodium sulphinide (prepared by adding the calculated amount of potassium sulphate to the barium salt in boiling water and filtering) and an alkaline solution of dimethyl sulphate were used.

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ATMOSPHERIC OZONE

BY HARRY N. HOLMES

The presence of ozone in the atmosphere is an unsettled problem. Schöne[1] and Ilosvray[2] considered it unproved and many authorities agree with them. Others believe with Schönbein, Houzeau, Andrews and many of the earlier investigators, that ozone is a normal constituent of pure air. None maintain that there is ever more than one part in seven hundred thousand[3] by volume. Haghurst and Pring,[4] who based

[1] Ber. d. chem. Ges., **13**, 1503. Z. anal. Chem., **33**, 137 (1894).
[2] Bull. soc. chim., **2**, 377. Ber. d. chem. Ges., **27**, 920 (1894); **28**, 2837 (1895).
[3] Houzeau: Compt. rend., **74**, 712.
[4] J. Chem. Soc., **97**, 868.

their estimate upon test reagents sent up by kites and balloons, state that below eight thousand feet the proportion of ozone is probably one in four billion, at very great heights, one in three million. Their method of determination, the oxidation of potassium iodide to potassium iodate, is, perhaps, open to question.[1]

Most observers estimate the proportion at about one part in one million by volume. D. Mendeléef[2] states: "The air of fields and forests always contains ozone."

The great work of Schönbein, whose sixty papers on ozone formed the first chapter in its history, entitles his opinion to consideration. Schöne, however, insisted that the tests obtained by Schönbein were due to the hydrogen peroxide of the air, to the trace of amyloid in the paper or to the ammonium carbonate always found in the atmosphere. Schönbein had used, among other tests, manganese sulphate papers, which were turned brown by ozone. Schöne found that traces of ammonium carbonate produced the same effect on manganese sulphate papers.

Houzeau, using potassium iodide-litmus paper in testing for atmospheric ozone, found the paper turned blue, because of the potassium hydroxide formed by the action of ozone on the potassium iodide. Schöne, however, contended that hydrogen peroxide produced the same result, and that the amyloid of the paper was a disturbing factor.

Schöne detected appreciable quantities of hydrogen peroxide in rain, snow and artificial dew, and insisted that the atmospheric ozone tests of his time were traceable to the hydrogen peroxide. He further stated that the odor he smelled when near a lightning flash was not the same as the odor of ozone, with which he felt quite familiar. It is interesting to note that Schöne could detect a pungent odor on the clothes of people entering a room from out of doors, especially in winter, but he felt sure it was not due to ozone. He admitted

[1] Thierry (Compt. rend., **134**, 460–3) found four times as much ozone at a height of two miles in the Alps as in Paris. He reported an amount of ozone greatly in excess of that given by Haghurst and Pring.

[2] "Principles of Chemistry," 1905.

that porous bodies, such as flannel, take up and retain some ozone.

Many theories have been advanced to explain the presence of this gas in the atmosphere. Besanez[1] states that evaporation of water invariably produces ozone and that this, rather than the silent electric discharge, forms the ozone in nature. Others consider that the rapid evaporation of water under the influence of sunlight produces hydrogen peroxide. The silent electric discharge and the action of ultraviolet light are held responsible by most writers of the present. Gerard[2] found warm, moist air very unfavorable to its formation. There is no question that the presence of moisture greatly reduces the yield of the silent discharge machines now used in the industries. Dust rapidly destroys ozone. Andrews[3] found that, when shaken in a flask with fine glass dust, the ozone is destroyed.

Carius[4] observed that ozone and ammonia gas form white fumes. He represented the reaction by the following equation:

$$2NH_3 + 4O_3 = NH_4NO_2 + H_2O_2 + 4O_2$$

This may have some bearing on the fact that rain during a thunder storm contains more nitrite and nitrate than ordinary rain. The destructive action of heat, moisture, dust and ammonia gas may possibly be aided by the slight solubility of ozone in water (about five parts in one thousand), as affirmed by Schöne,[5] Carius[6] and Houzeau.[7] Others contend that ozone is not soluble in water, but if it were rain would tend to wash it from the air.

Ladenburg and Lehman[8] have shown that ozone has a number of absorption bands scattered over a wide range of the spectrum. By a study of these ozone absorption bands in

[1] Ann. Chem. (Liebig), **181**, 232.
[2] Soc. Belg. Elec., Aug., 1909.
[3] Ann. Chem. (Liebig), **174**, 31.
[4] Ber. d. chem. Ges., **8**, 1481.
[5] *Ibid.*, **6**, 1224.
[6] *Ibid.*, **6**, 806.
[7] Ann. chim. phys., [4] **27**, 15.
[8] Ann. Physik, **21**, 305 (1906).

the solar spectrum, Angström[1] has detected the presence of what seems to be a very considerable amount of ozone, which must be in the upper air since it is not in the lower. Van Aubel[2] believes there is no question about the ozonizing action of short wave length. This action is best when the air is dry and cold—as in the upper strata.

Humphreys[3] expects to find considerable ozone in the higher air and but little near the earth. This he confirms by observations. "Only traces of ozone are found in the lower air and when there is much moisture and the temperature is high, the amount is vanishingly small." He states that an ozone blanket around the earth is much less permeable to earth radiations than to solar rays and therefore tends to make the terrestrial temperature more uniform. The effect of auroral displays, Humphreys affirms, must be the formation of an excess of ozone in the high latitudes and this would cause a greater diathermacy of the upper air at the equator, resulting in a comparatively low temperature and great altitude of the equatorial isothermal air. This hypothesis is sustained by the facts, for the isothermal air in the tropics is five kilometers higher and 15° C. colder than over the middle latitudes.

Humphrey's conclusions are:

1. An increase in sun spots appears certainly to be accompanied by a decrease in terrestrial temperature, at least in many places, fully twentyfold that which can be accounted for by the decrease in radiation from the spot areas alone.

2. It seems nearly certain that sun-spot maxima, whatever the value at such times of the solar constant, must lead to a decrease in the ultraviolet radiation that reaches the earth, and a corresponding decrease in the production, by this method, of ozone in the upper atmosphere.

3. The increase in the auroral discharges that accompany spot maxima tends to increase the amount of ozone, especially in the higher latitudes.

4. The change in the temperature of the earth, and all its train of consequences, from spot maximum to spot minimum,

[1] Ark. Mat. Astrom. Fysik, 1, 395 (1904).
[2] Compt. rend., 150, 96 (1910).
[3] Astrophysical J., 1910, 97.

is not necessarily dependent upon a change in the solar constant. It may depend largely, if not wholly, upon a change in the absorptive property of the atmosphere, caused, we believe, by a variation in the amount of ozone produced by ultraviolet radiation and by auroral discharges.

The whole question of the presence and detection of atmospheric ozone is complicated by the possible presence of two other oxidizing gases, hydrogen peroxide and nitrogen peroxide, which duplicate nearly all the ozone tests. The formation of nitrogen peroxide by the lightning flash is unquestioned. Berthelot[1] and Hautefeuille à Chappuis[2] report that it is formed to a slight extent by the silent discharge, though Briner and Durand[3] find no nitrogen peroxide formed by the silent discharge at low temperatures.

Theories of the formation of hydrogen peroxide in nature have already been mentioned. The exhaustive work of Schöne[4] on the presence of hydrogen peroxide in rain, snow and dew have a direct bearing on my own work, so some of Schöne's results will be given.

Schönbein[5] found hydrogen peroxide in rain and Struve[6] found it in both rain and snow.

Schöne used as reagents starch-potassium iodide with a little ferrous sulphate as a catalyst and a solution of guaiacum and malt extract. With these reagents he found from one part hydrogen peroxide in one million to one part in twenty-five million, or between 0.04 mg. and 1 mg. per liter of rain. The tests were made at Warsaw, Russia, and Schöne detected more hydrogen peroxide in the rain brought by certain winds than in others, also more in rain than in snow. The maximum, he observed, was in July, rapidly decreasing to December (the term of his experiments).

He found no hydrogen peroxide whatever in natural dew or frost, but by condensing the moisture of the air on a jar of ice, he was able to form artificial dew at any time. When

[1] Compt. rend., **92**, 87.
[2] *Ibid.*, **92**, 80, 134.
[3] *Ibid.*, **115**, 1272.
[4] Ber. d. chem. Ges., **7**, 1693.
[5] J. prakt. Chem., **106**, 272 (1868).
[6] *Z.* anal. Chem., **11**, 28 (1872).

this artificial dew was collected in the middle of a sunny day, the content of hydrogen peroxide was greatest—in July as much as 0.4 mg. per liter of dew. Collected at night or in the early morning or evening, none was detected. The rapid decrease in the hydrogen peroxide content during August, September, October and November indicates that the action of the sun's ultraviolet rays on the moist oxygen of the lower air was the probable cause of this formation.

Schöne calculated by formula from these tests with rain and artificial dew that the maximum content of hydrogen peroxide in air at Moscow was one part in thirty thousand million by volume. Obviously a direct test in air could not be secured.

EXPERIMENTAL

With a view to securing some approximate information as to the variation in the ozone content of pure air with the change in the seasons and during sudden atmospheric disturbances, I made a great number of starch-potassium iodide tests, some in June and July of 1911, but most of them in November, December, January and February. Of course this simple test is open to the old objection that any oxidizing gas will produce a blue color with the reagent. However, ozone, nitrogen peroxide and hydrogen peroxide are the only oxidizing gases found in normal air, and by frequent check tests, the hydrogen peroxide was absolutely eliminated and the share of nitrogen peroxide in the starch-potassium iodide blue proved to be small and infrequent. It is admitted that these check tests were all made in the winter but a study of Schöne's data will show a direct test for hydrogen peroxide in the air to be, to say the least, improbable.

To prove the absence of hydrogen peroxide, I made use of the wonderful catalytic effect of a little ferrous sulphate on the reaction between starch-potassium iodide and hydrogen peroxide. It accelerates the speed of this reaction to an astonishing degree.

For example, a solution of hydrogen peroxide so dilute that five minutes were required for the development of a blue color

with starch-potassium iodide turned a deep blue instantly when a drop of ferrous sulphate solution was added with the starch-potassium iodide. When two test papers were exposed, one wet with a starch-potassium iodide solution and the other the same with a little ferrous sulphate added, I expected the second to develop a blue first, if hydrogen peroxide was present. This opinion was justified on examining the air from a large commercial ozonizer. The paper treated with ferrous sulphate turned blue noticeably sooner than the other, showing that the ultraviolet glow acting on moist oxygen forms hydrogen peroxide as well as ozone.

This test, though frequently made in January and February, always failed to indicate any hydrogen peroxide in the atmosphere. To confirm these results, air was drawn through a solution of titanium sulphate for several hours on different occasions. Schöne[1] considers this an extremely delicate test but in no case did any yellow color appear.

To detect nitrogen peroxide, different tests were tried. The method of Andrews,[2] passing air through a tube heated well above 270° C. (which destroys all ozone but leaves nitrogen peroxide unchanged) and then over a starch-potassium iodide paper, failed to indicate any nitrogen peroxide. On other occasions, air was drawn for hours through a weak solution of pure sodium hydroxide, the solution acidified and tested by the Griess reagent (sulphanilic acid and α-naphthylamine) for nitrous acid. Generally no pink color could be seen and the test never indicated more than a very small amount of nitrogen peroxide.

Papers wet with a solution of benzidine in alcohol were frequently exposed. They yield a brown color with ozone, but blue with nitrogen peroxide and with hydrogen peroxide. As far as ozone is concerned, the test is not satisfactory, for the color of the dry benzidine is a red-brown and the distinction is hard to draw. But with hydrogen peroxide eliminated, any blue must indicate nitrogen peroxide. On a few stormy

[1] Z. anal. Chem., **9**, 41.
[2] P. Roy. Soc., **16**, 63.

days, a slight blue on the edge of the benzidine paper was observed, followed by an ozone brown.

The conclusion is that, at least in the winter, the development of the starch-potassium iodide blue is usually due to ozone alone, occasionally aided to a slight extent by nitrogen peroxide, but without any appreciable assistance from hydrogen peroxide.

Tetramethyl-*p*-phenylenediamine[1] was used to confirm the ozone test but it is open to the objection that nitrogen peroxide and hydrogen peroxide also turn it violet. The speed of the reaction is remarkable. The color frequently developed in a few seconds, almost instantly on the best ozone days, but in a room heated by hot air the test required minutes instead of seconds. Of course, most of the ozone had been destroyed when the air passed around the furnace.

An approximate idea of the relative quantites of ozone in the atmosphere at different times was obtained by noting the time of the appearance of the first trace of blue color developed in neutral starch-potassium iodide papers. Of course, it is not claimed that there is an exact quantitative relation between the time of the test and the amount of ozone in the air, but it can be roughly stated that the quicker the test, the greater the percentage of ozone. A large number of such simple tests with frequent checks on the presence of nitrogen peroxide and hydrogen peroxide must be of considerable value in indicating the effect of changing climatic conditions.

June and July were selected as typical summer months. At no time could tests with starch-potassium iodide be secured in less than seventy-five minutes, even on the finest June days, while the very hot weather of July, 1911, afforded no trace of ozone. At that time I did not try the more delicate tests used in the following winter. During a violent summer thunder storm, a test was obtained in fifteen minutes, yet so unstable was the ozone in the hot, moist air that twenty minutes later none could be detected.

[1] Wurster: Ber. d. chem. Ges., **19**, 3195.

Ozone Observations, November–February

Made at Richmond, Indiana

Date	Time of initial color Wet starch-KI paper	Temperature at time of test (Centigrade)	Weather conditions
Nov. 13	25 min.	Cold	Clear and bracing
Nov. (later) Several tests	Poor	Warm	Cloudy and depressing
Nov. (later) Several tests	45 min.	Cold	Bright and clear
Nov. 29	40 min.	0°	Snappy
Dec. 1	40 min.	0°	Cloudy and windy
" 4	45 min.	0°	Clear—no wind
" 7	30 min.	0°	Clear
" 10–25	Poor	Warm	Damp and depressing
" 27	30 min.	— 2°	Bracing, clear
" 29–30	Poor	Warm	
" 31	Fair test	Colder	"High" barometric area approaching
Jan. 1	50 min.	0°	Light breeze
" 2	Poor	Warm	Unsettled
" 3	40 min.	0°	Bracing
" 4	23 "	— 8°	Snow flurries
" 5	22 "	— 1°	Cloudy
" 8	40 "	— 8°	Light snow, "High" area passed day before
" 9	30 "	0°	Clear and windy
" 10	45 "	— 8°	
" 11	25 "	—14°	Snow on ground
" 12	17 "	—16°	Clear—in a "High" area
" 13	17 "	—16°	"High" area just passing to the East
" 14	Poor	— 7°	"Low" area approaching from the West
" 15	12 min.	—18°	West winds—"High" area approaching
" 16	11 min.	—22°	"High" area passed. Another in the S. E.
" 17	90 "	0°	In a "Low" area. Raining
" 18	45 "	5°	"High" area in S. E. Drizzling rain

(*Table continued*)

Date	Time of initial color Wet starch-KI paper.	Temperature at time of test (Centigrade)	Weather conditions
Jan. 19	24 min.	—11°	In a "High" area
" 22	Poor	5°	"Low" area on N. W. Snow flurry
" 23	60 min.	0°	"High" area far to the South
" 24	60 "	0°	"High" area approaching from the West
" 25	35 "	0°	
" 26	100 "	— 1°	"Low" area just passing
" 27 A. M.	60 "	— 4°	"High" area far to the North
" 27 P. M.	120 "	— 4°	"Low" area on West
" 29	180 "	— 4°	In a "Low" area
" 30	4 "	— 6°	In a "High" area
" 31	15 "	— 6°	Small "High" in South. "Low" in North
Feb. 1	15 "	— 2°	West wind
" 2	6 "	—11°	Strong "High" area near
" 3	15 "	—11°	"Low" area passed south of us
" 4	8 "	—20°	Strong "High" area came from N. W.
" 5	15 "	—18°	"High" areas in N. W. and S. W.
" 6	11 "	—18°	"High" area on the west
" 7	4.5 "	— 4°	"High" areas in N. W. and S. W. Snow flurries
" 8	4 "	—15°	Strong "High" near on N. W.
" 9	5 "	—12°	"High" area near on N.

On comparing the above records with the daily weather map it was evident that the maximum ozone content of pure country air was found only when "high" barometric areas were so near that the air from great heights rapidly flowed down to the observer. Owing to the greater extremes in temperature, the storm areas are more frequent and intense in winter, hence the quicker transfer of very high air to the surface. The low temperatures of winter are also more favor-

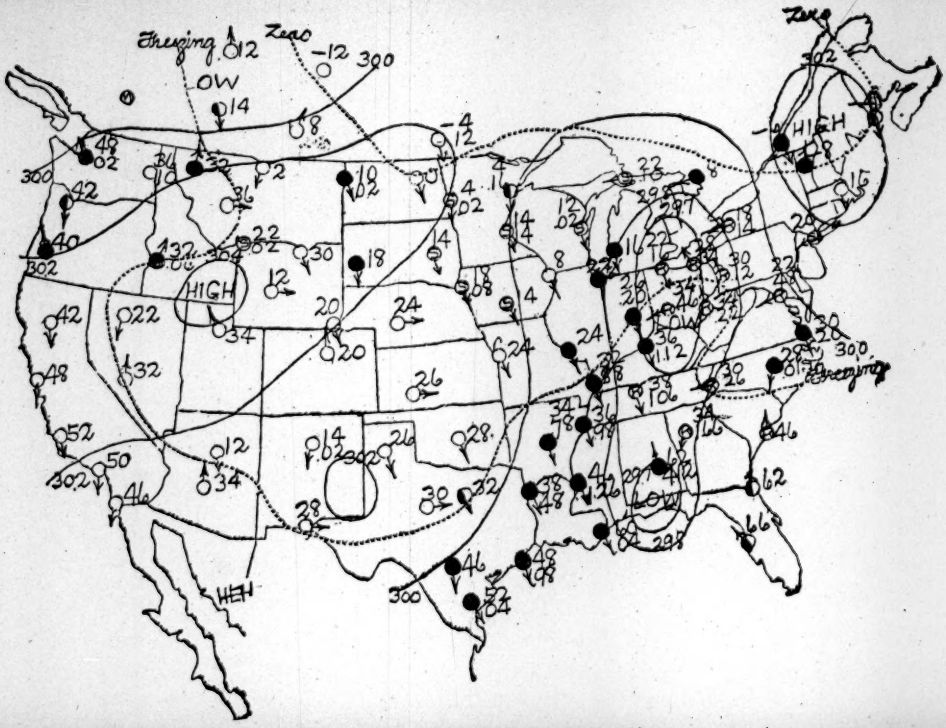

Fig. 1. Official Weather Map for January 29, 1912.

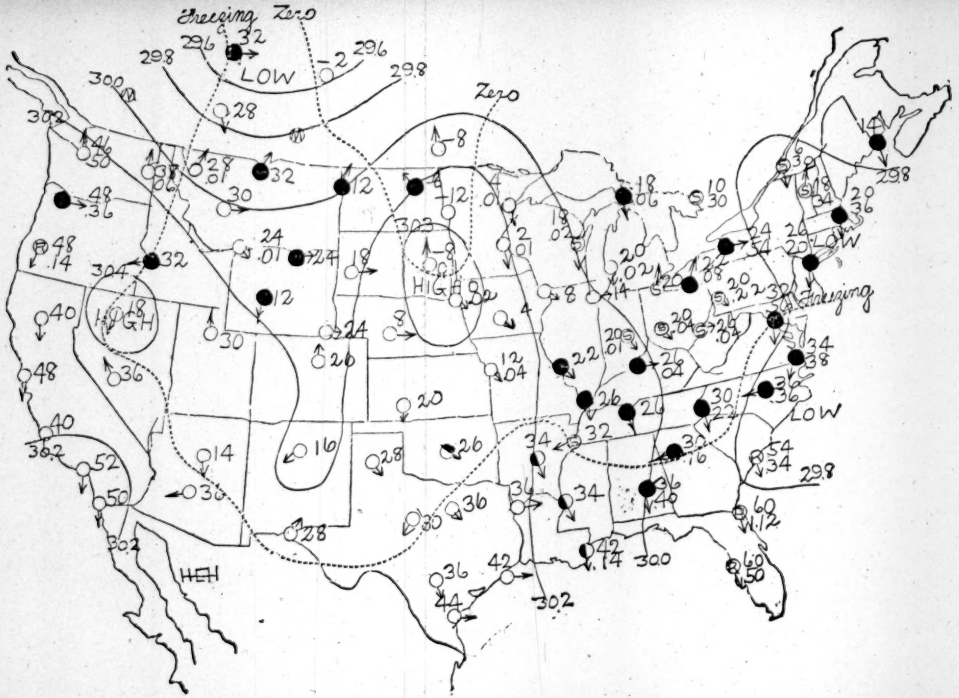

Fig. II. Official Weather Map for January 30, 1912.

able to the preservation of this ozone coming from the upper strata. The downward flow in summer is so slow and the temperature so high that very little ozone reaches the surface. Furthermore, in winter the ground is often covered with snow for a long period, an excellent check on the destructive action of dust.

Two weather maps, those of January 29th and 30th, 1912, will serve as typical illustrations. On January 29th, Richmond (on the eastern border of Indiana) was in the center of a "Low" area. A "High" area lay over the Rockies. The starch-potassium iodide tests required three hours for the initial blue color. The next day, January 30th, the "Low" had reached the Atlantic coast and Richmond was on the edge of a strong "High" just advanced from the Rockies. The best test of the year—a blue in four minutes—was obtained. Of course the air was clear and cold, yet the amount of ozone present was certainly not proportional to the temperature. Note that on this record occasion, the temperature was only 6° C. below zero, yet this test was far better than on another day when the extreme temperature was 34° C. below zero. Low temperatures help preserve the ozone, yet the strongest factor to be considered is the quick transference of extremely high air to the surface.

These results agree remarkably well with the theory of van Aubel, Humphreys and many others,[1] that our atmospheric ozone is for the most part formed by the action of the sun's ultraviolet rays on the upper air, which is much more permeable to these rays than the denser air near the earth. It is very probable that in our latitude the auroral discharge may be a factor. A further study of this subject is in progress.

CONCLUSIONS

1. The appreciable presence of ozone in winter air is proved by the starch-potassium iodide test. Hydrogen peroxide is present, if at all, in amounts too small to affect reagents much

[1] Henriet and Bonyssy (Compt. rend., **146**, 977) found that as the carbon dioxide in air increases the ozone decreases and *vice versa*. Their conclusion that this is explained by variations in the amount of upper air brought to the surface agrees with this paper.

more sensitive than this. Nitrogen peroxide is present occasionally but not always, and, at its maximum, in amounts that have little effect on starch-potassium iodide.

2. The maximum amount of ozone in the air is found when rapidly moving "High" barometric areas are so near the observer that quantities of air from great heights rapidly flow down to the earth.

3. The role of thunder storms in forming ozone is probably a minor one. We get most of our ozone from the great supply formed in the "Upper Laboratories" by the ultraviolet rays of sunlight acting on oxygen.

4. The ozone content of winter air is many times that of summer air.

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RICHMOND, INDIANA
February 14, 1912

A NEW BROMINE METHOD FOR THE DETERMINATION OF THYMOL, SALICYLATES AND SIMILAR COMPOUNDS[1]

BY ATHERTON SEIDELL

In a previous paper on methods for the determination of salicylates,[2] a large number of results were given from which it was concluded that neither the bromine method of Freyer,[3] which is essentially the well known Koppeschaar method, nor the iodine method of Messinger and Vortmann[4] gave satisfactory results when applied to salicylates. Nevertheless, it was thought desirable to give these methods a thorough trial with thymol, especially since results by the iodine method, as quoted by Messinger and Vortmann, indicated reasonable accuracy and, aside from the *p*-diazonitroaniline method of Riegler,[5] no other methods for the determination of thymol could be found in the literature.

[1] Presented at the meeting of the Division of Pharmaceutical Chemistry of the American Chemical Society, Washington, D. C., Dec. 28, 1911. A preliminary report of this method was published in the J. Wash. Acad. Sci. for Nov. 4, 1911.

[2] Seidell: J. Am. Chem. Soc., **31**, 1168 (1909).

[3] Chem. Ztg., **20**, 820 (1896).

[4] Ber. d. chem. Ges., **23**, 2753–55 (1890).

[5] Bul. Soc. Scii. Bucuresci, **8**, 51 (1899).

Before attempting to apply the bromine and iodine methods to thymol a few experiments were made to ascertain if an oxidation process could be devised for the quantitative determination of this compound. The action of potassium permanganate upon an alkaline solution of thymol was first tried and although a rapid decomposition of the permanganate, with production of manganese dioxide, occurred, it was soon ascertained that this product itself reacted with the permanganate and made it impossible to determine by the color of the solution when an excess of permanganate had been added. The filtrate from the black manganese dioxide precipitate never had the permanganate color but was a dirty green, and furthermore an additional amount of the precipitate separated from it on standing. Attempts to prevent this decomposing action of the manganese dioxide by adding a zinc salt were unsuccessful, even when tried in acid as well as alkaline solution. Other oxidizing agents, such as nitric acid, hydrogen peroxide, ferric sulphate, etc., were also found unsatisfactory.

The *iodine method*, although stated by Messinger and Vortmann to be applicable to thymol, was known from previous work on salicylates to have two serious defects, namely, the inaccuracies caused by the reasonable variation in the proportion of caustic alkali and those due to rather slight differences in the excess of standard iodine solution required for the titration. In the experiments with thymol it soon became apparent that these same defects are inherent in the method as applied to thymol. Analytical results confirming this conclusion are given in the following tables.

From the results in Table I it is seen that a moderate increase of the excess of iodine causes considerable increase in the apparent percentage of thymol. It should also be noted that with a still greater increase of excess of iodine the satisfactory flocculation of the precipitate fails, even in spite of the heating and shaking. During these experiments it was found that if the volume of solution was increased with distilled water to about 300 cc. or more, the iodothymol compound did not precipitate but simply colored the solution red and starch

Table I.—Determination of Thymol by the Method of Messinger and Vortmann

(Sample, 0.4242 gram thymol dissolved in 113.1 cc. 0.1 N NaOH and solution diluted to 250 cc.; mol. ratio thymol : NaOH = 1 : 4)

			Standard solution of			
Det. No.	Thymol solution cc.	Thymol Gram	Iodine cc	Thiosulphate (for one-half total filtrate) cc.	Calc. thymol found Gram	Calc. thymol Per cent.
1	40	0.06788	19.5	0.45	0.06787	100.0
2	40	0.06788	20.0	0.6	0.06862	100.9
3	40	0.06788	21.0	1.0	0.06936	102.2
4	40	0.06788	22.5	1.8	0.0690	101.6
5	40	0.06788	24.0	2.5	0.06934	102.2
6	40	0.06788	25.0	2.7	0.07155	105.4

(1 cc. I sol. = 0.012345 gram I, 1 cc. thiosulphate = 0.01249 gram I)

paste could be used as indicator for the end point of the direct titration with standard iodine solution. It was therefore apparent that by this simple change one of the two sources of error of the Messinger and Vortmann method could be obviated. A series of titrations was then made with this modification and the following results obtained.

Table II.—Determination of Thymol by the Modified Method of Messinger and Vortmann

(Sample, 0.3705 gram thymol dissolved in 98.79 cc. 0.1 N NaOH and solution diluted to 250 cc. with water; mol. ratio of thymol : NaOH = 1 : 4)

Det. No.	Thymol solution cc.	0.1 N H_2SO_4 or NaOH cc.	Resulting mol. ratio thymol: NaOH	Standard iodine solution cc.	Calc. thymol Gram	Calc. thymol Per cent.
1	30	2.97 (H_2SO_4)	1 : 3	11.5	0.04224	95.0
2	30	2.22 (H_2SO_4)	1 : 3.25	11.5	0.04224	95.0
3	30	1.49 (H_2SO_4)	1 : 3.50	11.7	0.04297	96.65
4	30	0.74 (H_2SO_4)	1 : 3.75	12.0	0.04407	99.13
5	30	0.0 (H_2SO_4)	1 : 4	12.1	0.04444	99.95
6	30	0.74 (NaOH)	1 : 4.25	12.25	0.04499	101.2
7	30	1.49 (NaOH)	1 : 4.50	12.4	0.04554	102.4
8	30	2.97 (NaOH)	1 : 5	12.5	0.04591	103.2

(1 cc. standard iodine solution = 0.01243 gram I)

These titrations show very strikingly that to obtain theoretical results there must be used exactly four molecules of alkali to each molecule of thymol, therefore requiring that the nearly exact quantity of thymol which it is desired to estimate be known in advance of the determination.

Since the allowable variation of excess of alkali was found to be so small when sodium hydroxide was used, it was thought that with a weaker alkali the permissible variation in its amount might be sufficient to allow of accurate determinations without a previous knowledge of the quantity of thymol to be determined. The following experiments (Table III) were, therefore, made.

Table III.—Determination of Thymol by the Modified Method of Messinger and Vortmann

(Sample, 0.8873 gram thymol dissolved in 118.0 cc. 0.1 N. NaOH and solution diluted to 200 cc. with water; mol. ratio thymol to NaOH = 1 : 2)

Det. No.	Thymol solution cc.	0.1 N alkali added cc.	Mol. ratio thymol: alkali		I sol. required (starch paste indicator) cc.	Calc. thymol Gram	Calc. thymol Per cent.
1	20	11.8 ($NaOH$)	1 :	4($NaOH$)	24.5	0.0890	100.2
2	20	11.8 (Na_2CO_3)	1 :	$2NaOH$ $2Na_2CO_3$	22.7	0.0824	92.9
3	20	17.7 (Na_2CO_3)	1 :	$2NaOH$ $3Na_2CO_3$	23.0	0.0835	94.1
4	20	35.4 (Na_2CO_3)	1 :	$2NaOH$ $6Na_2CO_3$	23.6	0.0857	96.6
5	20	118.0 (Na_2CO_3)	1 :	$2NaOH$ $20Na_2CO_3$	24.6	0.0893	100.7
6	20	354.0 (Na_2CO_3)	1 :	$2NaOH$ $60Na_2CO_3$	28.2	0.1024	115.3
7	20	35.4 (NH_4OH)	1 :	$2NaOH$ $6NH_4OH$	23.8	0.0864	97.0
8	20	118.0 (NH_4OH)	1 :	$2NaOH$ $20NH_4OH$	24.5	0.0890	100.2
9	20	excess $NaHCO_3$		.....	(end point uncertain)		

(1 cc. standard iodine solution = 0.012278 gram iodine)

The above experiments show that although the sodium hydroxide can be replaced by weaker alkalis no especial ad-

vantage is to be gained by so doing. Somewhat greater latitude in the absolute amount of alkali is allowable but there must still always be present the equivalent of 4 molecules of sodium hydroxide per molecule of thymol. These results indicate that this reaction might be used as a rough quantitative measure of the relative strengths of bases. Thus it is seen that 20 molecules of sodium carbonate or of ammonium hydroxide are the equivalent of 2 molecules of sodium hydroxide It was found that sodium bicarbonate, sodium acetate, potassium acetate, sodium tetraborate, sodium formate and several other salts permitted the reaction to proceed to a certain extent but the end point could not be ascertained with certainty. The results in several cases indicated that the reaction ended more or less completely with the formation of the monoiodothymol instead of the diiodo compound. A number of experiments in which these weak alkalis were used in connection with alcohol, to dissolve the thymol and, therefore, hasten the reaction, showed that in practically all cases iodoform was produced and could be detected by its odor. Experiments upon the use of more or less insoluble alkalis, such as calcium, magnesium, aluminium, etc., hydroxides, gave unsatisfactory results.

Attempts were finally made to control definitely the ratio of caustic alkali to thymol by simultaneously adding the standard iodine and sodium hydroxide solutions, thus insuring an end point at which both the solutions are present in equivalent molecular amounts, as is required by the ratio of 1 molecule of thymol to 4 of sodium hydroxide and to 4 of iodine. The titrations were made by placing in a large flask enough of the standard 0.1 N sodium hydroxide solution to dissolve the weighed sample of thymol, then adding an equal volume of 0.1 N iodine solution and continuing the alternate addition of alkali and iodine until the end point, as shown by starch paste, was reached. The first column of figures under NaOH in the following table shows the amount of alkali used to dissolve the thymol sample. The solutions were diluted with about 300 cc. of water in each case. The results obtained are given in the following table (No. IV):

Table IV.—Determination of Thymol by Simultaneous Introduction of Standard Alkali and Iodine

Det. No.	Wt. of thymol	0.1 N NaOH cc.	0.1 N iodine cc.	Thymol Gram	Thymol Per cent.
1	0.1233	20 + 10.8	30.8	0.1153	93.5
2	0.1346	20 + 15.0	35.0	0.1310	97.0
3	0.0892	20 + 3.9	23.9	0.0895	100.3
4	0.2155	30 + 23.7	53.7	0.2010	93.3
5	0.1435	20 + 16.4	36.4	0.1362	95.0

In the above titrations it was noted that although a sharp end point could be obtained, the results appeared to be influenced by the amount of sodium hydroxide used in the first place for dissolving the sample. When this quantity was near that which was theoretically required for the reaction, the results approached more nearly 100 per cent.

In a paper by Gardner and Hodgson,[1] upon the iodination of phenols and their iodometric estimation, the authors propose a modification of the Messinger and Vortmann iodine method which they claim gives concordant results for tannic acid and many other phenolic substances. It was found that in the case of some phenols one molecule of iodine was required for each hydroxyl group and in others quantitative, although empirical, results were obtained. The modification consists essentially in adding the standard iodine before the alkali and avoiding an excess of the latter. This is done by adding starch paste to the mixture of the tannic acid solution and excess of standard iodine and titrating with the alkali until the blue color disappears. Dilute acid is then added in sufficient excess and the unabsorbed iodine determined by means of standard thiosulphate. Results are given for a number of compounds but none are reported for thymol.

In one determination made by me upon thymol according to the method of Gardner and Hodgson, it was found that the blue color of the starch disappeared very slowly upon addition of 0.1 N sodium hydroxide, in fact it did not disappear completely even after the addition of a considerable excess of alkali. The determination could, therefore, not be finished in accordance with the proposed method.

[1] J. Chem. Soc., **95**, 1819–27 (1909).

During the past year, J. M. Wilkie has published[1] a paper on this subject in which the manner of substitution of the iodine in various phenols is discussed in detail and the view of Gardner and Hodgson that the substitution takes place in the hydroxyl group is criticized. It is found that in the case of β-naphthol a dilute aqueous solution, treated with 0.1 N iodine, yields monoiodo-β-naphthol, thus proving that the hydroxyl as such does not enter into the reaction. The author proposes a modification of the Messinger and Vortmann method which in principle is identical with that of Gardner and Hodgson; that is, suggests the addition of the iodine before the alkali. He, however, substitutes sodium carbonate for the hydroxide and omits the use of starch for avoiding too great an excess of alkali. Although Gardner and Hodgson report that salicylic acid absorbs an amount of iodine which does not correspond to a definite molecular ratio, Wilkie finds that by the details he proposes practically theoretical results corresponding to the formation of the triiodo compound are obtained. It would, therefore, appear that the procedure of Wilkie is the better, at least for salicylates. In view of my previous experiments upon salicylates it was considered desirable to make some determinations by the Wilkie method. A series of titrations upon sodium salicylate are given in Table V.

Table V.—Titration of Sodium Salicylate by Wilkie's Modification of the Messinger and Vortmann Method

Salicylate solution cc.	Sodium salicylate Gram	0.1 N iodine cc.	Approx. 0.1 N Na_2CO_3 cc.	0.5 N H_2SO_4 cc.	0.1 N thiosulphate cc.	0.1 N iodine consumed cc.	Calc. sodium salicylate Gram	Calc. sodium salicylate Per cent.
25	0.125	42.24	40	16	15.22	27.02	0.0721	57.7
25	0.125	52.80	50	20	20.36	32.44	0.0865	69.2
25	0.125	63.36	60	24	26.35	37.01	0.0987	79.0
25	0.125	73.92	70	28	32.84	41.09	0.1095	87.7
10	0.05	31.68	30	12	14.62	17.06	0.0458	90.8
10	0.05	42.24	40	16	23.48	18.76	0.0500	100.0
10	0.05	47.54	45	18	28.63	18.91	0.0505	100.9
10	0.05	52.80	50	20	33.90	18.9	0.0504	100.9
10	0.05	63.36	60	24	44.44	18.9	0.0505	100.9

[1] J. Chem. Soc. Ind., **30**, 398–402 (1911).

These results show that when sufficient excess of iodine and corresponding excess of alkali are used theoretical results are obtained. A considerable increase in the amount of excess of iodine does not, however, increase the percentage of salicylate. In making the titrations it was noted that the end point is ordinarily reached without preliminary color changes, hence, great care must be taken not to pass over it. Furthermore, the change is somewhat obscured by the presence of the reddish precipitate which gives the starch iodide a purple-blue color changing with excess of thiosulphate to the reddish color of the iodosalicylate precipitate. The accuracy of the single determination given by Wilkie is confirmed by the above results, and the method as applied to salicylates is considered to be the most satisfactory of any method so far proposed. Although Wilkie's modification of the Messinger and Vortmann method was found to apply satisfactorily to salicylates, the results obtained on thymol were variable and, furthermore, did not correspond to a definite molecular proportion, as will be seen from the following table (No. VI).

Table VI.—Determination of Thymol by Wilkie's Modification of the Messinger and Vortmann Method

(Sample, 1.2229 grams thymol dissolved in 81.32 cc. 0.1 N NaOH and solution diluted to 500 cc. with H_2O; mol. proportion thymol to NaOH = 1 : 1)

Thymol solution cc.	0.1 N iodine cc.	0.1 N Na_2CO_3 cc.	0.5 N H_2SO_4 cc.	0.1 N thiosulphate cc.	0.1 N I consumed cc.	Calc. thymol Gram	Thymol Per cent.
25	21.12	20	8	7.07	14.05	0.0527	86.2
25	42.24	40	16	28.51	13.73	0.0515	84.3
25	63.36	60	24	50.07	13.29	0.0499	81.6
			0.1 N NaOH				
25	42.24	10	4	28.51	13.73	0.0515	84.3
25	42.24	40	16	28.15	14.09	0.0528	86.5
25	42.24	80	32	28.03	14.21	0.0533	87.2

The end points of the titrations were fairly sharp although suspended material interfered with its detection to a certain extent. The blue color of the starch reappeared fairly rapidly. The calculation of the amount of thymol was made on the basis of 4 atoms of iodine per molecule of thymol. Although

the results do not indicate a reaction according to definite molecular proportions they are roughly quantitative even when the excess of iodine and alkali is varied very considerably. This is evidently a case similar to those placed by Gardner and Hodgson in the group with which the absorption of iodine takes place in empirical but (nearly) quantitative amount.

The Koppeschaar[1] *bromine method* for the determination of phenol has been claimed by Zdarek[2] to give satisfactory results with thymol. Detailed figures are however, given by him for only one determination and in this case the thymol found corresponded to 102.3 per cent.[3] The experiments recorded in the following table (No. VII) show that titration with the ordinary Koppeschaar bromate-bromide solution gives high and variable results, depending upon the amount of excess of standard bromate solution and the time allowed before titrating the excess of bromine.

Table VII.—Direct Bromate (Koppeschaar's) Titration of Thymol

(Sample, 1.2229 grams thymol + 81.32 cc. 0.1 N NaOH diluted to 500 cc. with water. The titrations were each made in a volume of solution of about 400 cc.)

0.1 N bromate cc.	Thymol cc.	Time[4]	0.1 N Thiosulphate cc.	0.1 N bromate consumed cc.	Calc.[5] thymol Gram	Calc. thymol Per cent.
50	25	10	32.6	17.4	0.0653	106.8
45	25	30	25.84	19.16	0.0719	117.6
40	25	10	21.93	18.1	0.0679	111.1
35	25	30	16.17	18.83	0.0707	115.6
30	25	10	13.0	17.0	0.0638	104.3
25	25	10	8.03	16.97	0.0637	104.2
20	25	10	3.41	16.59	0.0623	101.8
18	25	35	1.38	16.62	0.0624	102.1

[1] Z. anal. Chem., **15**, 233 (1876).

[2] *Ibid.*, **41**, 227–231 (1902).

[3] One determination of the solubility of thymol in water at 19°.4 made by Zdarek gave a value about 16 per cent. in excess of that read from the solubility curve which I have recently determined with the aid of the analytical method described in the following pages.

[4] This column shows the time in minutes allowed between acidifying the aqueous mixture with 10 cc. of hydrochloric acid (1 volume concentrated acid and 1 volume water) and adding the 5 cc. of 20 per cent. aqueous potassium iodide solution preparatory to titrating with standard thiosulphate.

[5] On basis of 4 atoms of bromine per molecule of thymol.

It would appear from these results that in aqueous solution the reaction goes somewhat beyond the formation of dibromothymol, that is, that the bromine acts too energetically. It is possible that by reducing sufficiently the excess of bromine and decreasing the time of its action on the thymol, theoretical results could be obtained. Experiments along this line were, however, not considered worth while since trustworthy accuracy could not be expected by such a plan.

Among the various modifications of the bromate method which suggested themselves, the most promising appeared to be an estimation of the hydrobromic acid formed in the reaction, instead of measuring the amount of bromine consumed, as hitherto exclusively done. Efforts to do this by adding bromine vapor to aqueous thymol solutions containing suspended calcium carbonate were unsuccessful, since a blank without the thymol showed the formation of considerable calcium bromide. Also the removal of the excess of added bromine, both by titration with alkali and shaking out with an immiscible solvent such as carbon bisulphide, were found unsatisfactory since in each case the subsequent titration of the hydrobromic acid with standard alkali gave high results. Attempts were then made to remove the excess of free bromine by adding potassium iodide and titrating with thiosulphate. After thus removing the free bromine it was found that the hydrobromic acid could be readily determined by the ordinary iodide-iodate reaction, according to which one atom of iodine for each molecule of strong acid is liberated as represented in the equation $5KI + KIO_3 + 6HBr = 6KBr + 6I + 3H_2O$. Since an excess of potassium iodide was already present it was simply necessary to add potassium iodate and continue the titration with the same standard thiosulphate, the difference between the first and second readings of thiosulphate corresponding to the hydrobromic acid. Although it was found that by this method of titration nearly theoretical results for thymol could be obtained, the greater the excess of bromine used to react with the thymol, the higher was the apparent percentage obtained. In fact, a survey of all the experiments showed that in practically every case high results

were obtained, thus indicating that bromine in ordinary aqueous solution acts too energetically upon thymol, just as had been found by the preliminary direct bromate titrations of thymol (Table VII) and as previously reported for salicylates.[1] As mentioned above, carbon bisulphide had been used to remove the excess of bromine in several cases, being introduced, however, after the addition of the bromine vapor. In several subsequent determinations, in which the carbon bisulphide was added before the bromination, it was accidentally noted that the results in these cases did not exceed the theoretical 100 per cent. A series of determinations was then made to test this point in more detail, with results shown in the following table (No. VIII).

Table VIII.—The Bromine-Hydrobromic Acid Method for the Determination of Thymol

Det. No.	Thymol Gram	CS_2 cc.	Standard thiosulphate for Free Br cc.	Standard thiosulphate for HBr cc.	Diff.	Calc. thymol Per cent.
1	0.2405	none	9.3	41.6	32.3	102.3
2	0.1109	none	6.3	21.6	15.3	105.1
3	0.1522	none	5.4	26.0	21.6	108.1
4	0.0	5	9.15	9.2	0.05	0.0
5	0.5502	5	10.3	77.9	67.6	100.4
6	0.1178	5	8.3	22.8	14.5	100.5
7	0.3905	5	25.9	73.0	47.1	98.6
8	0.3124	5	25.4	63.6	38.2	99.9
9	0.2229	5	34.8	62.0	27.2	99.7

(Factor for standard thiosulphate: Det. Nos. 1–3, readings thiosulphate × 1.014 = 0.1 N exactly; Det. Nos. 4–9, readings thiosulphate × 1.0887 = 0.1 N exactly)

These determinations show that when carbon bisulphide is used, it is practically impossible to make the reaction go beyond the formation of the dibromothymol and obtain results higher than the theoretical. The carbon bisulphide (or any other immiscible solvent) evidently serves to remove the thymol from the aqueous layer and, therefore, permits the reaction to take place mainly, if not entirely, in the organic solvent. It is evidently the influence of this medium which

[1] Seidell: J. Am. Chem. Soc., 31, 1168 (1909).

controls the too great reactivity of the bromine. As might be expected, this controlling action of the immiscible solvent is also exerted in the ordinary direct bromate (Koppeschaar's) titration of thymol, as will be seen by comparing the results given in Table VII with the following:

Table IX.—Direct Bromate (Koppeschaar's) Titration of Thymol in Presence of Carbon Bisulphide

(Sample, 1.2229 grams thymol + 81.3 cc. 0.1 N NaOH diluted to 500 cc. with water)

Thymol solution cc.	0.1 N bromate cc.	CS_2 cc.	HCl (1 vol. conc. HCl + 1 vol. H_2O) cc.	0.1 N thio-sulphate cc.	0.1 N bromate used cc.	Calc. thymol Per cent.
25	25	5	10	9.04	15.96	98.0
25	40	5	10	23.84	16.16	99.2
20	40	5	5	27.07	12.93	99.2
25	50	5	20	32.46	17.54	107.6

Although the presence of carbon bisulphide undoubtedly prevents the reaction going beyond the formation of the dibromothymol the results which were obtained in the above and a great many other series were quite variable and this modification of the bromate method was, therefore, finally abandoned. The cause of the irregularities appeared to be that during the time which must be allowed (0.5 hour) for the completion of the reaction between the thymol and bromine some of the excess of bromine (added as standard bromate) disappeared, and since this loss would be calculated as thymol, the results varied according to the time allowed before the back titration. That the fault lay with the bromate solution itself was proved by titrating given amounts of standard bromate after allowing each to stand with carbon bisulphide and excess of acid for varying lengths of time; thus, 25 cc. 0.1 N bromate plus 25 cc. N hydrochloric acid plus 5 cc. carbon bisulphide treated immediately with 5 cc. of 20 per cent. potassium iodide solution required 25.1 cc. 0.1 N thiosulphate; when allowed to stand 0.5 hour before adding potassium iodide, only 23.6 cc. of 0.1 N thiosulphate was required; after 1 hour, 22.8 cc.; and after 2 hours, 22.1 cc. Thus, in

2 hours 12 per cent. of the total free bromine disappeared and this, calculated as thymol, would correspond to 112 per cent. of the actual amount present. The use of Koppeschaar's solution, even with the modification involving the addition of carbon bisulphide, is, therefore, not applicable to the determination of thymol.

Since the experiment just described showed that carbon bisulphide is slowly acted upon by bromine, in aqueous solution, it followed that an error due to this cause must also exist in the newly devised thymol method. Such was found to be the case and experiments were therefore undertaken to find a means of overcoming the difficulty. It was soon found that carbon tetrachloride was not acted upon by bromine to an appreciable extent. Its use, however, in place of carbon bisulphide resulted unsatisfactorily at first, since the reaction of the bromine upon the thymol appeared to be incomplete and furthermore the end point of the first thiosulphate titration was quite uncertain. After many experiments, it was found that the use of smaller quantities of carbon tetrachloride permitted the reaction to go to an end, and the uncertainties of the thiosulphate titration were almost entirely overcome by adding carbon bisulphide just previous to beginning the titration.

The details of the method as finally elaborated are as follows: The weighed sample of 0.1 to 0.5 gram of thymol is placed in a 300 cc. glass-stoppered bottle with 1 to 2 cc. carbon tetrachloride and 100 cc. water. Bromine vapor is then poured into the mixture until the color, after thorough shaking, shows that considerable excess of bromine is present. After one-half hour, 5 cc. of carbon bisulphide and immediately thereafter 5 cc. of aqueous 20 per cent. potassium iodide solution are added and the liberated iodine corresponding to the free excess of bromine is titrated with 0.1 N thiosulphate; an additional amount of potassium iodide solution is added and if no further liberation of iodine occurs the reading on the burette is taken. Five cc. of aqueous 2 per cent. potassium iodate solution are then added and after thorough shaking the titration with thiosulphate is continued until the iodine

color is just discharged for the second time. The completion of the reaction is tested by a further addition of potassium iodide and iodate solutions. The difference between the first (which should be from about 5 to 15 cc. 0.1 N thiosulphate) and the second readings corresponds to the hydrobromic acid formed by the action of bromine on the thymol. The calculation is made on the basis of two molecules of hydrobromic acid per molecule of thymol; 1 cc. 0.1 N thiosulphate is, therefore, equal to 0.0075056 gram thymol.

A series of determinations gave the following results:

Table X.—The Determination of Thymol by the New Bromine-Hydrobromic Acid Method

Gram thymol	Time (min.)	cc. thiosulphate for excess Br	cc. thiosulphate for HBr	Diff.	Gram thymol found	Per cent. thymol
0.0774	30	18.0	42.0	24.0	0.0778	100.6
0.0643	30	21.4	41.2	19.8	0.0642	99.8
0.0916	30	19.6	48.0	28.4	0.0920	100.5
0.0757	30	17.8	41.0	23.2	0.0752	99.4
0.1029	60	20.2	52.0	31.8	0.1030	100.2
0.1549	120	18.4	66.1	47.7	0.1546	99.8
0.0560	30	18.2	35.5	17.3	0.0561	100.1
0.0296	30	18.5	27.7	9.2	0.0298	100.8
0.0301	60	18.85	28.2	9.35	0.0303	100.7
0.0295	60	12.35	21.35	9.00	0.0292	99.0

(Factor for standard thiosulphate: readings $\times$ 0.8640 = 0.05 n exactly.)

Numerous experiments upon several samples of thymol have shown that the above procedure gives results well within a limit of 1 per cent. of the theoretical. Mr. R. M. Chapin, of the Bureau of Animal Industry, Department of Agriculture, kindly made a number of independent determinations with the following results:

Det No.	Thymol Taken Gram	Thymol Found Gram
1	0.2504	0.2509
2	0.5215	0.5225
3	0.2492	0.2507
4	0.2028	0.2035

He suggests, as a matter of convenience, that an approximately 5 per cent. solution of bromine in carbon bisulphide be used

in place of bromine vapor; the free acid in such a solution must, however, be removed before use by shaking out thoroughly with water. Experiments by Mr. Chapin on *benzophenol* indicated that by allowing sufficient time for bromination almost theoretical results, calculated on the basis of three substituted bromine atoms instead of two as in the case of thymol, are obtained.

Titrations made by the above method upon a sample of *sodium salicylate* gave the following results:

Wt. of sample	Standard thiosulphate for Excess Br cc.	Standard thiosulphate for HBr cc.	Diff.	Calc. $C_6H_4OHCOONa$ Per cent.
0.1257	15.25	30.0	14.75	100.4
0.1710	16.0	35.65	19.65	100.7

(Factor for thiosulphate: readings × 1.094 = 0.1 N exactly)

The calculation was made on the basis of the formation of the dibromosalicylate. The end points of the titrations are sharp and no other precautions than those mentioned in the description of the method as applied to thymol are necessary.

In this connection, attention should be called to the fact that in my previous work on salicylates[1] the formation of dibromosalicylate under somewhat different conditions was mentioned, namely, by the action of bromine in very strong hydrochloric acid solutions. It would, therefore, appear that in the present case the control of the too vigorous action of the bromine has been accomplished in a much more satisfactory manner. The method of titrating in presence of a great excess of hydrochloric acid is admittedly tedious and was only suggested as a possible improvement over the ordinary bromate titration in which the rather uncertain tribromophenolbrom compound is formed. In a recent paper by Autenrieth and Beuttel[2] the authors have attempted to determine phenols, salicylic acid, etc., by the gravimetric estimation of the tribromophenolbrom compound. They obtained results on phenol varying between 96.3 and 97.5 per cent.

[1] J. Am. Chem. Soc., **31**, 1168 (1909).
[2] Arch. Pharm., **248**, 112 (1910).

and on salicylic acid varying between 95.9 and 96.7 per cent. These determinations confirm in general my own experiments, which were referred to in my paper (p. 1177) as follows: "(*b*) *Bromate Gravimetric Methods.*—The compound formed by passing bromine vapor through a dilute aqueous solution of a salicylate was found to be too soluble for its successful removal and furthermore it could not be dried to constant weight." A careful examination of the paper by Autenrieth and Beuttel indicates that the same conclusion may be applied to their results, if by quantitative accuracy is understood determinations which are at least within 1 per cent. of the theoretical.

In regard to *sodium benzoate* and *benzoic acid* it was found that no bromination occurs under the conditions of the determination. In fact, on adding bromine vapor to an aqueous solution of sodium benzoate the yellow color of the bromine is immediately imparted to the solution. In the case of benzoic acid, however, although no bromine is absorbed the addition of potassium iodide and iodate to the aqueous solution of benzoic acid causes a liberation of iodine due to the slow reaction of the free benzoic acid according to the equation already given for the action of the iodide-iodate mixture on hydrobromic acid. A satisfactory determination of benzoic acid cannot be made by this method, however, on account of the slowness with which the end point is reached. It has been found that even weaker acids than benzoic acid react to a greater or less extent with the iodide-iodate mixture and, therefore, cannot be determined by the present method. It was found that even sodium benzoate liberates iodine after some time from the iodide-iodate mixture. This point then constitutes one of the limitations of the method and must be kept in mind in any attempts to apply it to other compounds than thymol.

In view of the fact that under the conditions of the present method, bromine reacts with salicylates but not benzoates, it would appear that this procedure offers a simple method for the quantitative determination of salicylates in the presence of benzoates or of mixtures of salicylic and benzoic acids. In the latter case it is only necessary to convert the mixture

into sodium salts by titration with standard sodium hydroxide solution and then make the determination of the salicylate as above described, the benzoate being obtained by difference. The end point of the titration of salicylate should be determined quickly, however, since on standing the benzoate would react very slowly with the excess of iodide-iodate solutions, as mentioned above.

Experiments with *phenyl salicylate* and *methyl salicylate* indicated that in the case of the former compound no substitution of bromine occurred, while with the latter, although considerable bromine was absorbed and sharp end points for the titrations were obtained, the results did not show that the reaction proceeded according to a definite molecular ratio. Several determinations indicated that carbon bisulphide prevents the absorption of as much bromine as in the absence of this solvent.

In the case of *resorcinol* it was found that on account of its great solubility in water, carbon bisulphide was not able to withdraw enough of it to prevent the reaction taking place almost entirely in the aqueous layer. The bromination, therefore, proceded further than would be expected in the organic solvent, the result being that on adding potassium iodide to remove the excess of free bromine a reaction between this reagent and the bromoresorcinol compound occurred and the end point with thiosulphate could not be obtained. This experiment indicates another limitation of the new method, namely, that the compound to be determined should be relatively much more soluble in the organic solvent than in water.

It was found that in brominating *acetanilide* in the presence of carbon bisulphide the aqueous layer did not show an excess of bromine until some time after the carbon bisulphide layer was colored a deep red-brown. It was therefore necessary to add a very considerable excess of bromine to reach the desired end point. The titrations were very sharp but, as seen from the following table of results, considerable variation was obtained. Since, however, a difference of 0.1 cc. thiosulphate corresponds to 0.5–1.0 per cent., the variations can not be considered excessive.

Table XI.—Determination of Acetanilide by the Bromine-Hydrobromic Acid Method

Acetanilide Gram	Thiosulphate for Free Br cc.	Thiosulphate for HBr cc.	Diff.	Calc. $CH_3CONH.C_6H_5$ Per cent.
0.1496	7.25	16.95	9.4	101.2
0.1712	27.1	37.9	10.8	101.1
0.2934	41.35	60.0	18.65	102.3

(Factor for thiosulphate: readings × 1.192 = 0.1 N exactly)

Attention should be called to the fact that in this method of brominating acetanilide in the presence of carbon bisulphide only one bromine atom is substituted in the ring, whereas by the bromination in aqueous solution[1] of acetanilide saponified by strong hydrochloric acid the tribromo compound is formed. The present method does not require the previous saponification of acetanilide but, on the other hand, does not yield as closely agreeing results.

With *phenacetin* it was found practically impossible to carry the bromination to the end point; furthermore, the titrations with thiosulphate were unsatisfactory since the carbon bisulphide layer assumed a dirty yellow instead of clear white color after disappearance of the free iodine.

In the case of β-naphthol it was found that only fairly accurate results could be obtained since, as in the case of acetanilide, quite small variations in the amount of thiosulphate corresponded to relatively large differences in the per cent. of β-naphthol. It was observed that, in brominating, a very large amount of bromine must be added to produce the yellow color in the aqueous layer. Although the substitution of this bromine in the β-naphthol evidently occurred, the product formed was readily decomposed by potassium iodide since in all cases the liberated iodine appeared to be much more than would result from the amount of excess of free bromine added, and, furthermore, the calculations of the absorbed bromine corresponded to the substitution of only one bromine atom in the ring. As will be seen from the table of results (Table XII) the thiosulphate required for the free

[1] Seidell: J. Am. Chem. Soc., **29**, 1091 (1907).

bromine was in every case from two to three times that corresponding to the hydrobromic acid, showing that at least two bromine atoms had been taken up in the first place, and on addition of excess of potassium iodide, one of these again liberated.

Table XII.—Determination of β-Naphthol by the Bromine-Hydrobromic Acid Method

β-Naphthol Gram.	Standard thiosulphate for Br cc.	Standard thiosulphate for HBr cc.	Diff.	Calc. $C_{10}H_7OH$ Per cent.
0.0971	16.8	22.5	5.7	100.9
0.2160	32.1	44.9	12.8	101.8
0.1950	27.0	38.7	11.7	103.0
0.1476	26.7	35.6	8.9	103.6
0.4041	47.0	71.3	24.3	100.9

(Factor for thiosulphate: readings × 1.192 = 0.1 N exactly)

In addition to the above enumerated compounds many others which are included in the U. S. Pharmacopoeia were tried, but for the most part, with unsatisfactory results. Of these the following should be mentioned: eugenol, vanillin, eucalyptol, antipyrine, sodium phenolsulphonate, camphor, menthol, camphoric acid and naphthalene.

HYGIENIC LABORATORY
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REVIEWS

PORTLAND CEMENT. Its Composition, Raw Materials, Manufacture, Testing and Analysis. By RICHARD K. MEADE, M.S., General Manager of the Tidewater Portland Cement Co. Second edition. Easton, Pa.: The Chemical Publishing Co. 1911. pp. x + 512. Price, $3.50.

The remarkable, we may almost say miraculous, growth of the Portland cement manufacture in this country and abroad for the past twenty years has drawn general attention to this industry. No other industry has shown such rapid development and no other material, with the exception of iron, has found so widespread use. Ten years ago Mr. Richard K. Meade, an instructor in Lafayette College, which is located in the then heart of the cement industry, became interested in the chemistry of Portland cement. A search through the

literature of chemical technology failed to reveal any definite information on this subject. On inquiry of cement manufacturers and their chemists, it was found that each was pursuing general schemes of analysis for the purpose of regulating the variations of raw material and finished product, but there had been no effort made to standardize these methods and, in fact, there seemed to be some hesitancy about the interchange of ideas and improvements as the tendency then was to make the process of manufacture as secret as possible. Mr. Meade then decided to gather what information he could and in 1902 published a little book entitled "The Chemical and Physical Examination of Portland Cement." This book was more in the nature of a manual and was intended as a guide for young chemists who were then being called in large numbers to conduct the routine analyses in the factory laboratories. This book explained how cement might be analyzed and tested but gave no hint as to the process and detail of manufacture.

A few years later there was published by the Ohio Geological Survey a bulletin on the "Manufacture of Hydraulic Cements." This bulletin was received with open arms by the industry, as its author, Dr. A. V. Bleininger, had made an exhaustive treatment of the subject and had supplemented his description with numerous records of original work. Seeing the reception which this book was accorded and perceiving that the demand for a book of this kind was greater than the Ohio State Survey was able to supply, Mr. Meade began the task of compiling from the sources then available a book of like nature. Accordingly, in 1906, he published his first edition of "Portland Cement," which he not only dedicated to the chemist, but also to the engineer, by adding much material on the manufacture of this product. While the contents of this book was largely a compilation of data and material which from time to time had appeared in the trade literature, yet its thorough treatment and convenient form assured it a welcome not only to the cement industry but also to others who were anxious to know just what this remarkable material was and by what methods it was made.

Since its publication the author has had considerable experience, not alone in the chemical end of manufacture, but also in the mechanical field, and in order to bring his work up to date and correct some of the statements on which he had received more light, he published a second edition in 1911, which is the subject of this review.

In the present edition the entire text has been revised and much new matter has been added, especially on manufacture,

so that the book is about evenly divided between technology and the chemistry and physical testing of cement. About 70 new illustrations have been added, many of them excellent half-tones showing machinery and actual installations.

A valuable feature has been added in the form of description of the raw materials and types of machinery used by a number of the best plants in this country and a list of references shows one where may be found descriptions of practically every cement plant in the United States.

Perhaps too much prominence has been given to descriptions of plants with which the author has been connected, but this is only natural, as one can best describe that with which one is most familiar. For the same reason greater prominence has been given to certain types of machines and this must be taken into consideration lest the reader mistake these for being the most efficient.

The latest results on the investigation of the constitution of Portland cement have been recorded and the author is to be commended for advocating so strongly the importance of finer grinding of both the clinker and raw materials. Mention is made of the efforts directed towards determining the "flour" or active cementitious material in commercial cement but the author errs in believing that anything can be learned from the sieving of cement, as the residue left on the finest wire sieve so far made (300-mesh) reveals nothing but very finely divided, vitreous clinker.

The new chapter on "Investigation of Materials for the Manufacture of Portland Cement" is of some value but is deserving of more thorough treatment, and because of its importance should be included in the chapter headed "Raw Materials."

On the whole, the book is one which should be found in the library of every manufacturer, whether his position be that of Manager, Superintendent or Chemist, for it contains much of interest and value to all, and in a convenient form can be found items for reference which otherwise would require considerable search. The book is neatly and substantially bound, the type and cut work excellent and the price reasonable.

CLARENCE N. WILEY.

DIE ZUSTANDSGLEICHUNG. Rede gehalten am 12 Dec., 1910, in Stockholm bei Empfang des Nobelpreises für Physik. Von PROF. DR. J. D VAN DER WAALS. Leipzig: Akademische Verlagsgesellschaft. M. b. H. 1911. pp. 24.

The lecture discusses the subject under four heads:

1. How van der Waals discovered the equation.

2. The present status of the equation in the mind of its author.

3. How, in the last four years, he has sought to explain the differences between the results of experiment and of his equation.

4. How he has sought to explain the behavior of binary and ternary mixtures by means of his equation.

It is always interesting to learn how any great investigation was begun: what incited it, and through what stages it passed.

Van der Waals tells us in this lecture that the first incentive to take up the work which has consumed the best energy of his life was obtained from the work of Clausius published in 1857 on "Die Art der Bewegung welche wir Wärme nennen."

In this paper Clausius showed how the law of Boyle can be easily derived, by assuming that a gas consists of material points which move with high velocity.

The lecture is such a concise and condensed account of a large investigation that to give any really adequate account of it would be to give the lecture.

H. C. J.

TRAITÉ DE CHIMIE GÉNÉRALE. Par W. NERNST, Professeur à l'Université et Directeur de l'Institut de Chimie Physique de l'Université de Berlin. Ouvrage traduit sur la 6e édition allemande par A. CORVISY, Professeur Agrégé des Sciences Physiques au Lycée Gay-Lussac, Professeur Suppléant à l'École de Médecine et de Pharmacie de Limoges. Deuxième Partie: Transformations de la Matière et de l'Énergie. Paris: Librairie Scientifique A. Hermann et Fils. 1912. pp. 422. Price, Fr. 10.

This is the second and final part of Prof. Corvisy's translation of Nernst's book.[1] The only criticism which can be made of this book is the omission of a subject index; otherwise, Prof. Corvisy's work seems to have been done excellently.

C. A. R.

CONTEMPORARY CHEMISTRY. A Survey of the Present State, Methods and Tendencies of Chemical Science. By E. E. FOURNIER D'ALBE, B.Sc., A.R.C.S., M.R.I.A. New York: D. Van Nostrand Company. 1911. pp. xvi + 180. Price, $1.25.

This is a very readable account, in semipopular style, of the present state of chemistry, written from the standpoint of the physicist. The subjects taken up are the Molecule, States of Aggregation, Optical Chemistry, the Theory of Solutions, Osmotic Pressure, Affinity, Valency, Chemistry and Electricity, Chemical Analysis, Crystallization, Carbon Compounds, Chemistry and Life, the Chemistry of Metals, Industrial Chemistry, the Atomic Theory, Radio-activity and the Chemistry of the Future. While some inaccurate

[1] See THIS JOURNAL, **46**, 313.

statements are made, the book, as a whole, makes interesting reading and will serve excellently to while away an idle hour. The similes and metaphors scattered through it are refreshing, to say the least, as, for example, the following: " 'It is really impossible,' says Sir William Ramsay, 'to manipulate quantities of the order of a few tenths of a milligram.' Yet a tenth of a milligram, even of the heaviest element, contains at least half a trillion (0.5×10^{18}) atoms. As chemistry, properly so-called, deals only with tangible substances, which can be 'manipulated,' it follows that all purely chemical methods work with aggregates of units numbering over a million times the number of individuals in the whole human race. Imagine the teeming millions of China and India, the hordes of Africa, and the civilized populations of Europe and America lumped together and multiplied a millionfold—men, women, and children—and the whole aggregate subjected to some supra-chemical analysis. What would be the result? The 'substance' would turn out to be an 'element.' No amount of interaction with other masses of a different species would split it up into two or more independent substances. Hence the efforts of the supra-chemist would be necessarily confined to an investigation of its physical properties, and its ultimate structure. And, firstly, there would be the three states of aggregation. In the 'liquid' state of humanity, the 'atoms' would constantly remain within each other's sphere of influence, within hail or signaling distance. They would exhibit a good deal of cohesion and of what we have lately come to call 'steric hindrance.' When scattered beyond that distance the human units or atoms would act more or less independently, but always retain a tendency to form binary groups or higher 'molecular' associations. A tightly packed crowd would resemble a viscous and rather amorphous solid. A tendency to crystallization would be shown in the slow deposit of an orderly audience at a place of entertainment. At the 'absolute zero of temperature,' illustrated by the grave-yard, there would be evidence of crystallization in the quadratic system. And even the phenomena of Lehmann's 'living crystals' would be paralleled by an army on the march."

C. A. R.

Recherches Récentes sur les Faciès des Cristaux. Conférence faite le 25 janvier, 1911, par M. P. Gaubert, Docteur ès Sciences, Assistant de Minéralogie au Muséum National d'Histoire Naturelle. Publication de la Société de Chimie Physique. Paris: Librairie Scientifique A. Hermann et Fils. 1911. pp. 33. Price, Fr. 2.

In this pamphlet the author summarizes certain of his fun-

damental notions as to the inception and progress of crystal growth. Special emphasis is laid on the rôle of foreign substances in modifying the external shape of crystalline bodies. Most of the experimental data have been previously published in greater detail by the author or other investigators. A number of references are given to the literature of the subject.

The significance of currents in the mother liquor, of "fundamental faces," of absorption of coloring matter and other impurities contained in the mother liquor ("adsorption" of some authors) and of spherulitic and helicoidal structures will hardly be regarded in precisely the same light by all crystallographers. However, this discussion cannot fail to contribute toward an explanation of these interesting phenomena.

A. C. G.

FERTILITY AND FERTILIZER HINTS. By JAMES EDWARD HALLIGAN, Chemist in Charge, Louisiana Experiment Station. Easton, Pa.: The Chemical Publishing Co. 1911. pp. vii + 155.

This book is an abridged edition of the author's "Soil Fertility and Fertilizers." It deals with fertility, manures, chemical composition and properties of fertilizer materials, lime, gypsum, trade value of fertilizers and home mixtures. Comparatively little space is devoted to the practical application of fertilizers and the fertilizers suited to different crops. G. S. F.

DIE ZERESINFABRIKATION. Von INGENIEUR-CHEMIKER DR. PHIL. BELA LACH, technischer Konsulent für die Oel- und Fettbrauche in Wien. Mit 49 in den Text gedruckten Abbildungen. Monographien über chemisch-technische Fabrikations-Methoden. Halle a. S.: Druck und Verlag von Wilhelm Knapp. 1911. pp. x + 207. Price, M. 9.60.

The author calls attention to the fact that previous works on this subject contain really little that is technically valuable, and deal with old methods of working. He hopes this volume will supply the deficiency of books on the subject. It is handled in a large way—the occurrence, method of working, purification and testing of ozokerite being discussed in the first part.

The second part deals with the purified product, ceresene, and describes the bleaching with acid, animal charcoal and clays: its filtration, the presence of the wax and its extraction by solvents. Finally its adjuncts, paraffin and other waxes, its uses for candles, honeycombs, wax paper, etc., and methods of testing are thoroughly described.

The ambition of the author seems to have been realized, and the treatise may be cordially recommended. A. H. GILL.

QUANTITATIVE CHEMICAL ANALYSIS, Adapted for Use in the Laboratories of Schools and Colleges. By FRANK CLOWES, D.Sc., London, Emeritus Professor of Chemistry in the University College, Nottingham, and J. BERNARD COLEMAN, A.R.C.Sc., Dublin, Head of the Chemical Department in the Southwestern Polytechnic, Chelsea, S. W. Ninth edition. Philadelphia: P. Blakiston's Son & Co. 1911. pp. xxiv + 565. Price, $3.50.

The eighth edition of this excellent book was reviewed in THIS JOURNAL.[1] The present edition contains scarcely any new matter except the determination of sulphur in fuel oils.

E. B. SPEAR.

A BRIEF LABORATORY GUIDE FOR QUALITATIVE ANALYSIS. By ARTHUR E. HILL, PH.D., Associate Professor of Analytical Chemistry, New York University. Easton, Pa.: The Chemical Publishing Co. 1911. pp. vi + 80.

As the title indicates, this book is intended as a laboratory guide for a very elementary course in qualitative analysis. On reading the preface, one gets the impression that the author's principal reason for writing the book was that he could not find another book sufficiently elementary for his purpose—a half-year's course of six hours a week—and a reading of the text confirms this impression. An examination of the book fails to reveal any particular merit to commend it to the teacher of qualitative analysis.

B. F. L.

TRAITÉ COMPLET D'ANALYSE CHIMIQUE APPLIQUÉE AUX ESSAIS INDUSTRIELS. Par. J POST, Professeur Honoraire à l'Université de Goettingue, et B. NEUMANN, Professeur à la Technische Hochschule de Darmstadt, avec la collaboration de nombreux chimistes et spécialistes. Deuxième édition française, entièrement refondue. Traduite d'après la troisième édition allemande et augmentée de nombreuses additions par G. CHENU, ING. E. P. C., Licencié ès Sciences, et M. PELLET, ING. I.N.A., Licencié ès Sciences. Tome troisième, premier fascicule. Engrais Commerciaux, Amendements et Fumiers, Terre Arable et Produits Agricoles. Air. Huiles Essentielles. Cuir et Matières Tannantes. Colle. Tabac. Caoutchouc et Gutta-percha. Matières Explosives et Allumettes. Avec 56 figures dans le texte. Paris: Librairie Scientifique A. Hermann et Fils. 1912. pp. 468. Price, Fr. 15.

This is the French edition of the book reviewed in THIS JOURNAL,[2] dealing with fertilizers, soils, air, essential oils, tanning materials, leather, glue, tobacco, gutta percha, rubber, explosives and matches.

It is, however, more than a translation, additional methods, peculiarly French, being given, and it may unhesitatingly be recommended to those who prefer this language to German.

A. H. GILL.

[1] Vol. 44, page 199 (1910).
[2] Vol. 41, p. 448.